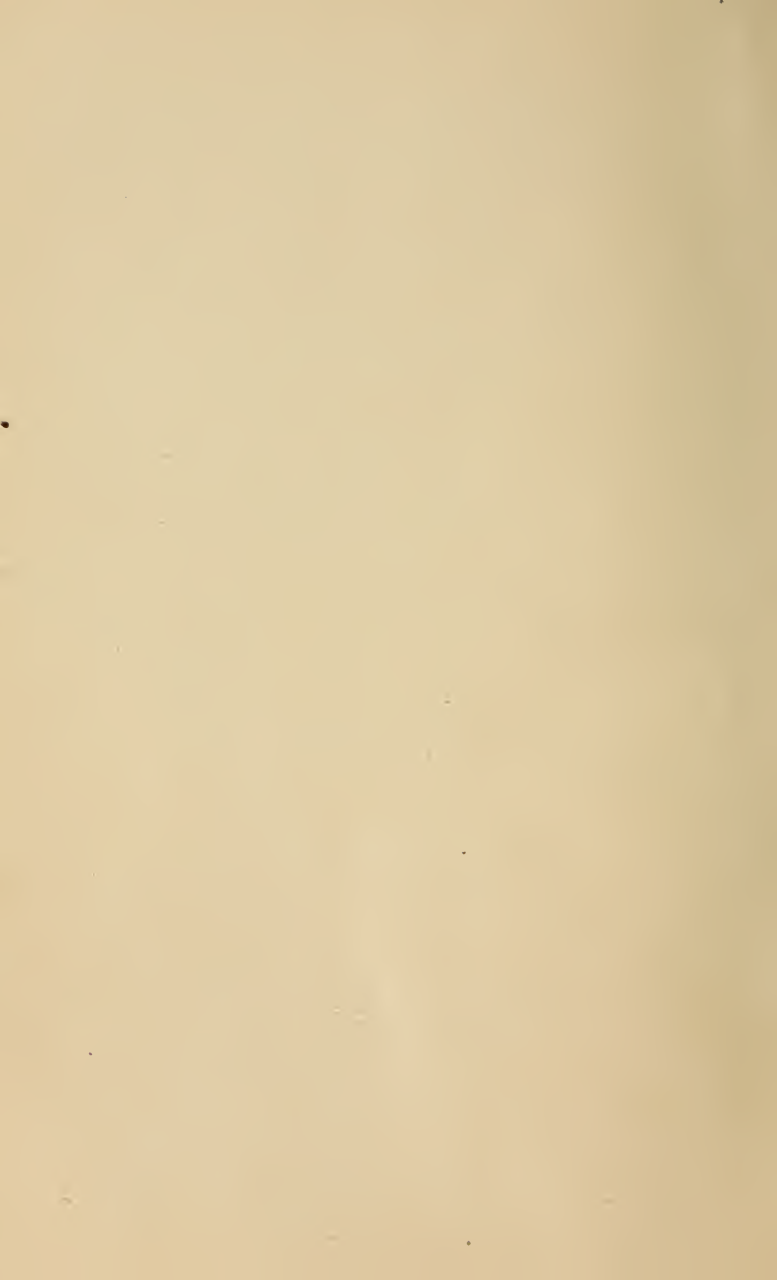


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AN ELEMENTARY COURSE
IN
INORGANIC PHARMACEUTICAL
AND MEDICAL CHEMISTRY.

DESIGNED ESPECIALLY FOR STUDENTS
OF PHARMACY AND MEDICINE.

BY

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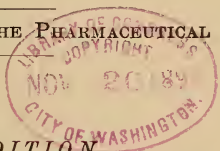
REVISED, AND REPRINTED FROM THE PHARMACEUTICAL RECORD.

FIRST EDITION.

FIRST THOUSAND.

NEW YORK:
JOHN WILEY & SONS,
53 EAST TENTH STREET.

1894.



53196-21

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ROBERT DRUMMOND, ELECTROTYPYER AND PRINTER, NEW YORK.

PREFACE.

THE Elementary Course in Pharmaceutical and Medical Chemistry was first published in the *Pharmaceutical Record* during the year 1892. Its original purpose was to direct and guide those pharmaceutical and medical students who had not the advantages of the personal direction of a teacher in their chemical studies. The course admirably carried out this purpose with its early readers, and its publication in book-form is in response to requests from many of these readers, and from teachers of chemistry. It was the author's intention to thoroughly revise the course, to eliminate much and to add more; but the many duties incumbent on a busy life have not permitted more than a partial revision. The method of presentation is the result of the author's experience as a teacher of pharmaceutical chemistry. The work is presented in such form that no part of it is complete without the other; it is necessary to study the book from beginning to end to follow its logical sequence. Repetition is adopted purposely; the statements which are repeated, perhaps frequently, are those which beginners do not always grasp the first time they are made. Repetition in an elementary course such as this is essential; it has the advantage which lies in the presentation of the same facts in so many different forms

of language that the reader will find at least one form understandable. In a few instances where it seemed well, the language of others is quoted to make clearer the point under discussion. Facts in chemical philosophy are scattered through the book and are given where they seemed to fit best and where their application is most appropriate. The course is graded in a manner which requires that it be studied page by page. Its plan is to gradually lead up to the comprehension of the chemistry of the inorganic salts official in the United States Pharmacopœia. The explicitness and detail of the treatment of the subjects up to the metals and their salts render brevity thereafter possible, in that the study of the salts requires essentially only the application of the principles laid down in the first half of the course. Where repetition is useless it is not employed; so it has been possible, to an extent, to study the properties of salts of the same metal collectively.

The author is gratefully appreciative of the aid rendered by those who have favored him with valuable suggestions, and to those who have so kindly assisted in reading and correcting the proof.

He desires also to express his gratitude for the favor with which the course was received during its first publication, and he indulges the hope that in its present form it will be still more worthy of the confidence of pharmaceutical and medical students and teachers.

MINNESOTA STATE UNIVERSITY,
MINNEAPOLIS, Nov. 1894.

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A COURSE IN ELEMENTARY INORGANIC PHARMACEUTICAL AND MEDICAL CHEMISTRY.

INTRODUCTORY.

STUDENTS in any branch of chemistry must all begin at the same starting-point. They must necessarily first acquire the knowledge of the fundamentals, the elements of the science, upon which, when once their own, they can construct or build any division of the science. So the student in pharmaceutical chemistry must first acquaint himself with the principles of chemistry without reference to the application of chemistry to pharmacy. When he has accomplished that, the principles may then be applied to pharmacy.

Molecular Forces.—In order to intelligently understand the subject it will be necessary to deviate a little and enter the field of physics, to get enough knowledge from that to better comprehend the meaning of chemistry.

The universe, the sum of all created things, is made up of what is termed matter, which material assumes three different forms, the solid, liquid, and gaseous. All things are in one of these three conditions, and the condition in which any given portion of matter exists is dependent upon the predominance of one or more of the natural forces. A given mass of matter is made up of *little particles* which are held together by the force called *cohesion*—that is, the particles of matter through the agency of this

force of *attraction* are made to unite to form a mass. If this force of cohesion is very strong, we have bodies which are very hard; is this force less strong, the bodies are soft or liquid. When this attractive force is altogether absent, and is replaced by another force having just the contrary tendencies—that is, to repel the particles from each other—we have gases as a result. This latter force is called *repulsion*.

It must be remembered that these forces act only on and between the small particles mentioned, which we will henceforth call *molecules*. A molecule is therefore the smallest particle of any kind of matter which will retain the characteristics or nature of the mass or body of which it is a part, and on and between which particles the forces of attraction and repulsion act. The latter two forces are often spoken of as *molecular forces*, because they concern themselves exclusively with molecules. The science which concerns itself with the molecular forces is termed physics or natural philosophy. Physics in a limited sense studies the nature of, and the phenomena observed in, bodies in which there is no change in composition or construction. Chemistry, on the other hand, comprehends only the *internal structure*, as it were, of the molecules. When we subdivide a given mass or body *physically* until it becomes impossible to further subdivide it without affecting its construction or internal structure, we arrive at what we call a molecule, and at this point we have reached the limit of physics, the sharp line which separates physics from chemistry. It is not impossible to further subdivide the molecule, but when we do so we destroy its internal structure and enter thereby the realm of chemistry. The molecule may be decomposed, resolved, into still smaller

particles, to which we apply the term "*atoms*." Atoms, therefore, make *molecules*, and molecules *mass*. We have learned that the molecules in forming mass by aggregating are held together by the force of attraction, and that this force is restricted to molecules. Now, the force which unites atoms to form molecules is a very different one from the molecular, and is only exerted among or between atoms.

The Atomic Force.—When a molecule has been resolved into its constituent atoms the process of subdivision has been carried to its utmost extent. The resolution of molecules into atoms usually gives us several kinds of atoms, and, because the atoms represent the ultimate division of matter, when we have different *kinds of atoms* we must have different *kinds of matter*. Chemists have found that there are between 65 and 70 different kinds of atoms, and consequently there must be as many different ultimate kinds of matter. Such an ultimate kind of matter, which, when its molecule is resolved into atoms, yields only one kind of atoms, is termed an *element*, or a simple substance. These 65 or 70 simple substances unite with each other in various proportions, and form the innumerable diversity of substances and bodies in the universe. This union is in obedience to the *chemical* or *atomic force*, or *chemical affinity*, or *chemism*, and takes place only between atoms. Apparently molecules unite with each other *chemically* at times; the union is, however, between the atoms constituting the molecules. There is as sharp a distinction between the molecular and atomic forces as there is between a molecule and an atom. When molecules unite to form a mass the latter has all the properties which the molecules had, or *vice versa*, but when atoms unite, the resulting molecules have no properties in common with the atoms which

formed them. In other words, when atoms unite with one another they lose their identity, their individuality, and the resulting molecules are totally different in all respects. For instance, if we unite a molecule of iron with another molecule of iron, we get a mass of iron, which still retains all the properties of the iron; but if we unite an atom of iron with an atom of iodine, we get a molecule of iodide of iron, which is totally different from either the iron or the iodine. In the latter union the chemical force exerts itself, in the former the molecular.

The student will thus far remember:

That atoms are the smallest particles of matter which will enter into chemical union, and

That this chemical union destroys the identity of the kind of matter acted upon, forming bodies altogether different from those entering into their composition.

That the chemical force is that force which induces and governs that kind of union.

That molecules are the smallest particles of matter which can retain the properties of the mass of which they are a part, and

That they are held together by the molecular force.

Physics relates to the study of bodies in which there is no change in composition.

Chemistry studies the phenomena observed in bodies in which there is a change in structure, and the law inducing the change.

The Elements.

THE 65 or 70 simple bodies or elements in an innumerable variety of combinations—chemical combinations—make up the sum of matter of the earth and all upon it.

In thus uniting to form this diversity of bodies, the elements behave principally in one of two ways (these ways will be alluded to later, when the student will have become better fitted to understand them), which ways afford us a convenient means for subdivision into two primary groups—the metals and non-metals. Each of these is for convenience further subdivided into smaller groups, the elements of each group having some property or properties in common, which makes their classification desirable.

The names of the elements are derived from the Greek and Latin, and usually indicate something of the nature of the element. For convenience and brevity symbols are employed in place of the whole names, these symbols being shorthand for the longer Greek or Latin names. Of the whole number of elements a few more than half are important enough to merit mention here, the remainder being of rare occurrence. The following represents a convenient classification of the elements into groups and sub-groups:

NON-METALS.

<i>Names.</i>	<i>Symbols.</i>	
Oxygen.....	O	} Oxygen group.
Sulphur.....	S	
Selenium.....	Se	
Tellurium.....	Te	
Chlorine.....	Cl	} Halogen group.
Bromine.....	Br	
Iodine.....	I	
Fluorine.....	Fl	
Phosphorus.....	P	}
Carbon.....	C	
Silicon.....	Si	
Boron.....	Bo	
Nitrogen.....	N	
Arsenic.....	As	} Metalloids.
Antimony (Stibium).....	Sb	
Bismuth.....	Bi	

METALS.

<i>Names.</i>	<i>Symbols.</i>	
Potassium (Kalium).....	K	} Alkalies.
Sodium (Natrium).....	Na	
Lithium.....	Li	
Ammonium.....	(NH ₄)	
Calcium.....	Ca	} Alkaline earths.
Barium.....	Ba	
Strontium.....	Sr	
Magnesium.....	Mg	} Magnesium group.
Zinc.....	Zn	
Cadmium.....	Cd	
Silver (Argentum).....	Ag	} Copper group.
Mercury (Hydrargyrum).....	Hg	
Copper (Cuprum).....	Cu	
Iron (Ferrum).....	Fe	} Iron group.
Manganese.....	Mn	
Cobalt....	Co	
Nickel.....	Ni	
Chromium.....	Cr	} Chromium group.
Molybdenum.....	Mo	
Antimony (Stibium).....	Sb	} Bismuth group.
Bismuth.....	Bi	
Platinum.....	Pt	} Platinum group.
Gold (Aurum).....	Au	
Aluminum.....	Al	
Lead (Plumbum).....	Pb	
Tin (Stannum).....	Sn	
Hydrogen.....	H	

The above, it must be remembered, are not all the elements known, nor are those mentioned in any one group all that may belong to that group; they are the more important ones. It will be seen that antimony and bismuth are found in both primary divisions—these and the metalloids will be considered further on.

The student at this point need only remember the names and symbols of the elements and the two primary divisions.

Physical Properties of the Elements.

The elements, as elements, exist in three forms; some are solid, some liquid and some gaseous. Mercury and bromine are liquid. Hydrogen, oxygen, chlorine, nitrogen and fluorine are gases, all the rest solids. The condition in which an element, or for that matter any compound, exists, is dependent upon the prevailing temperature. Many solids can be made to become liquid by the application of heat, and liquids to become gaseous. By the withdrawal of heat gases can be liquefied and liquids solidified. Oxygen can be liquefied by pressure and cold. Mercury will become solid, just as water will, if the temperature is low enough. On the other hand, mercury and water and other elements or compounds can be changed into gases by heat, in which condition they will remain as long as the temperature is not lowered.

Most of the elements are insoluble, odorless and tasteless. Sodium and potassium are soluble in water, uniting chemically with the water. Chlorine and bromine have very characteristic odors.

Nearly all have considerable chemical affinity, for which reason so few are found in nature in the elementary condition. A few exceptions are nitrogen, gold and copper. Sometimes an element occurs in more than one form, as carbon, with which we are familiar in the form of the diamond, graphite and charcoal; and oxygen, which sometimes assumes a nature which makes ozone of it. These conditions are termed *allotropic*, and are probably due to the different arrangement of atoms in the molecules. Two atoms of oxygen make a molecule of oxygen, but if the molecule contains three atoms of oxygen we call it ozone.

In this case the molecules contain different numbers of atoms, but when the number of atoms in the molecule is the same, the difference is in the arrangement of the atoms in the molecule. While some elements are *crystallized*—that is, assuming a definite geometrical outline—others are *amorphous*, just the contrary—that is, without any regular shape.

The metallic elements usually have a lustre, a color from the white silver to the bluish-gray lead, excepting copper, which is red, and gold, yellow. They differ considerably in their relative weights, hydrogen being the lightest and platinum the heaviest. The metals are, as a rule, heavy—platinum being $21\frac{1}{2}$ times as heavy as water. Lithium, sodium, and potassium are lighter than water.

Chemical Properties of the Elements.

ATOMIC WEIGHTS AND THE LAW OF DEFINITE CHEMICAL PROPORTION.—In studying the chemical behavior of the elements we enter upon the field of chemistry proper, and concern ourselves almost wholly with the atomic force, the force binding together the atoms to form compounds.

The atomic force, or *chemical affinity*, as it is frequently called, exerts itself at the point of contact of two or more atoms—that is, if the elements are capable of forming a compound; for it must be understood that chemical affinity does not exert itself upon or between *any* two or more atoms. This chemical union does not take place in a haphazard way; on the contrary, certain definite laws govern it. If we were to unite hydrogen and oxygen, we would find that the union takes place between *definite quantities* of the elements, and that if either were in excess of

its definite quantity, so much of either would remain over as either was in excess of its definite quantity. For illustration, let us observe closely the proportions in which hydrogen and oxygen will unite. Both are gases, but we may nevertheless *weigh* them, and in so doing, which is necessary, we find that hydrogen weighs less than oxygen does, bulk for bulk. If we are careful to take an equal *volume* of each gas, the hydrogen will be found to weigh just one-sixteenth as much as the oxygen does; or, on the other hand, that oxygen is 16 times heavier than hydrogen. Let the student bear this in mind for the present. The weights of *equal volumes* of the two gases may have been found to have been of hydrogen 10, and of oxygen 160 grains. If these two equal volumes were made to unite chemically (by touching the mixture with a lighted match), it would be found that all of the hydrogen, but only half the oxygen, entered into chemical union; in other words, that the 10 grains of the hydrogen united with 80 grains of the oxygen, and that the other 80 grains of oxygen remained unaffected. This experiment might be repeated with any desired quantity of the two gases, but always with the result that they would unite in the proportion of 1 of hydrogen and 8 of oxygen *by weight*, or 2 of hydrogen and 1 of oxygen *by volume*, and that if either be in excess of these proportions, that excess will not be affected, but will remain unchanged. The product of this combination is water. In like manner we could induce chemical union between hydrogen and chlorine, and the definite chemical proportions would be 1 and $35\frac{1}{2}$, respectively, *by weight*, or 1 and 1 *by volume*. This product would be hydrochloric acid.

In the case of hydrogen and sulphur in the gaseous con-

dition, the proportions *by weight* would be, of hydrogen 1 and sulphur 16, or *by volume*, 2 of hydrogen and 1 of sulphur.

Chemists have assumed that all elements in the gaseous condition contain in equal volumes, under like conditions, an equal number of molecules, and that molecules of elements usually consist of two atoms, and that, therefore, equal volumes of most elements in the gaseous condition contain an equal number of atoms. The weights of atoms must, according to this, differ to the extent that equal volumes of gases differ in weight. Hydrogen is the lightest gas, and if equal volumes of oxygen, chlorine and sulphur are, respectively, 16, 35.5 and 32 times heavier than hydrogen, and if we give the hydrogen atom the value of 1, it is obvious that the weights of the atoms of oxygen, chlorine and sulphur must necessarily be 16, 35.5 and 32. These figures, therefore, represent the *atomic weights* of these elements. It must be understood that these weights are *relative*. These figures denote not only the atomic weights of the elements, but they denote also the *proportions* in which they will unite with one or more hydrogen atoms. Oxygen will unite in the proportion of 16 with 2 of hydrogen, chlorine of 35.5 with 1 of hydrogen, and sulphur of 32 with 2 of hydrogen, for the reasons stated and others that will be stated further on.

In this wise the atomic weights of all elements have been obtained, in comparing their weights to that of hydrogen, either directly or indirectly. Another method will be spoken of a little further on.

LAW OF MULTIPLE PROPORTIONS.—The student must bear in mind also *that the proportions in which elements unite with hydrogen are the proportions in which they will*

unite with one another. Oxygen will therefore unite with sulphur in the proportion of 16 to 32. The atomic weight of carbon is found to be 12, of phosphorus 31, of nitrogen 14. Carbon, phosphorus or nitrogen would therefore unite with 16 parts of oxygen in the proportions expressed by the numbers expressing their atomic weights. We find, though, that these proportions—i.e., those expressed by the atomic weights—are not the *only* ones in which these elements will unite. If, in the case of carbon and oxygen, 32 of oxygen, twice 16, were taken, chemical combination would ensue, with the formation of a definite compound. Nitrogen and oxygen will unite not only in the proportion of 14 and 16, but in such a way that multiple proportions of the oxygen with a constant proportion of nitrogen will form definite compounds. The known combinations of nitrogen and oxygen may serve to illustrate:

	By weight.
Nitrogen monoxide.....	{ Nitrogen.....28 parts. Oxygen.....16 parts.
Nitrogen dioxide.....	{ Nitrogen.....28 parts. Oxygen.....32 parts.
Nitrogen trioxide.....	{ Nitrogen.....28 parts. Oxygen.....48 parts.
Nitrogen tetroxide.....	{ Nitrogen.....28 parts. Oxygen.....64 parts.
Nitrogen pentoxide.....	{ Nitrogen.....28 parts. Oxygen.....80 parts.

The prefixes mono, di, tri, tetra, and penta (oxide) here mean, respectively, one, two, three, four and five times one oxygen. It will be seen that the proportion of nitrogen (28, twice 14) is constant, while the quantity of oxygen increases by 16 in the ratio of one, two, three, four and five times 16.

Carbon and oxygen unite in these proportions:

Carbon 12, and oxygen 16, form carbon monoxide.

Carbon 12, and oxygen 32, form carbon dioxide.

This evidences the important law *that some elements unite in more than one proportion, and that the resulting definite compounds contain, to a constant proportion of one of the elements, multiple proportions of the other.*

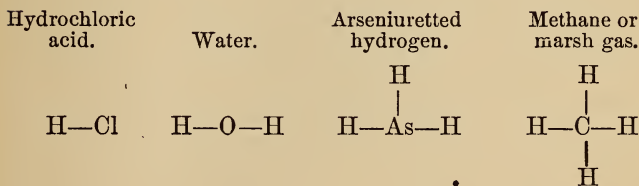
Carbon monoxide *always* contains certain fixed and definite proportions of carbon and oxygen; it cannot be carbon monoxide if it does not. This is true also of the carbon dioxide. All definite compounds consist of the same elements in the same proportions, according to our first law, and according to the second law, if two elements unite in more than one proportion, they unite in multiples of that proportion, and never in intermediate proportions.

Quantivalence.

We have learned that hydrogen was taken as a standard in comparing the weights of atoms of the elements. This is not its only function as a standard; it is also employed in establishing the *quantivalence* or value of atoms in replacing other atoms in combination. All atoms have value, and these values are constant in most atoms. We assume that all atoms have *bonds* with which they unite with each other, and that the *number* of bonds fixes the value of the atom in combining power. So hydrogen and chlorine unite atom for atom. So do hydrogen and iodine. These atoms have equal values. Hydrogen and oxygen unite in the proportion of two atoms of the former to one of the latter; three of hydrogen combine with one of arsenic; four of hydrogen combine with one of carbon. The oxy-

gen atom has twice the value of hydrogen, arsenic thrice, and carbon four times the value.

Oxygen, having twice the value of hydrogen or of any element uniting with hydrogen atom for atom, is said to have two bonds, each one of which combines with the one bond which each hydrogen atom has. Arsenic, because it requires three of hydrogen for combination, is said to have three bonds, and carbon, for similar reason, to have four bonds. The number of hydrogen atoms the atom of another element requires for combination (in case of possibility of combination) determines the number of bonds that atom has:



The value of an atom is not ascertained alone by studying the number of hydrogen atoms it requires for combination, but also by observing the number of other atoms it can replace in combination. In the above $\text{H}-\text{O}-\text{H}$ one of the hydrogen atoms can be *replaced* by one atom of potassium or one of sodium, hence the sodium and potassium atoms have the same value which the hydrogen atom has. If the oxygen were replaced by sulphur, $\text{H}-\text{S}-\text{H}$, the sulphur would have two bonds, because it replaced the oxygen, which needed two atoms of hydrogen for combination. The two hydrogen atoms might be replaced by one of calcium, $\text{Ca}=\text{O}$, therefore calcium would have two bonds. Any atom that could replace the calcium would

be twice the value of the hydrogen atom in combining power, for the same reason. The value of atoms might be better understood by assuming that the hydrogen atom is a quarter of a dollar, the *standard*, oxygen half a dollar piece, and carbon a dollar. To equal the dollar (carbon) with quarters (hydrogen), four of the latter would be required; to equal the half dollar (oxygen) with quarters (hydrogen), two of the latter would be needed, and to equal the dollar (carbon) with half dollars (oxygen), two would be required.

Special terms are used to designate the value of atoms. Those having the value of hydrogen are termed *monads*, or they are said to be *monatomic* (*mono*, one), or *univalent* (*unus*, one); those combining with or replacing two of hydrogen, *diads* (*di*, two) or *diatomic*, or *bivalent* (*bis*, two); those having thrice the value of the hydrogen atom are termed *triads* (*tres*, three), *triatomic* or *trivalent*; the next in order are *quadrads* or *tetrads* (*quadra*, four), *quadrivalent*; *pentads* (*penta*, five), or *pentatomic*; *hexads* (six); *septads* (seven), etc.

Atomicity and equivalency are terms synonymous with quantivalence and valence, but the term quantivalence is to be preferred.

The subjoined table gives the names of the elements, their symbols, quantivalence and atomic weights. The student will notice that many of the elements are classed in more than one column. Those elements have more than one value—they have as many as the table gives them. When an atom has more than one value, the values will always be expressed by even numbers or by uneven numbers, never by both. Thus nitrogen is triad, pentad and sometimes septod, but it never has two, four or six

bonds. In cases where the atomic weight is expressed by fractions, the student may remember the nearest whole number.

Elements—Quantivalence, Symbols, Atomic Weights.

Monads.	Diads.	Triads.	Tetrads.	Pentads.	Hexads.	Septads.
Hydrogen, H, 1	Oxygen, O, 16 Sulphur, S, 32	Nitrogen, N, 14.01	Carbon, C, 12	N P	S Fe	Cl Br
Chlorine, Cl, 35.37	Calcium, Ca, 39.91	Phosphorus, P, 30.96	Silicon, Si, 28.3	Cl Br	Al Cr	I
Bromine, Br, 79.76	Magnesium, Mg, 24.3	Gold, Au, 196.7	Platinum, Pt, 194.3	I As	Co Ni	
Iodine, I, 126.53	Copper, Cu, 63.18	Arsenic, As, 74.9	Fe Al	Sb Bi	Mn	
Fluorine, Fl, 19	Zinc, Zn, 65.1 Strontium, Sr, 87.3	Antimony, Sb, 119.6	Mn Cr			
Sodium, Na, 23	Cadmium, Cd, 111.5	Bismuth, Bi, 208.9	Co Ni			
Potassium, K, 39.03	Mercury, Hg, 199.8	Cl Br	Ca Ba			
Rubidium, Rb, 85.20	Tin, Sn, 118.8	I	S Pb			
Silver, Ag, 107.66	Barium, Ba, 136.9					
Caesium, Cs, 132.7	Lead, Pb, 206.4					
	Iron, Fe, 55.88					
	Aluminium, Al, 27.04					
	Manganese, Mn, 54.08					
	Chromium, Cr, 52					
	Cobalt, Co, 58.6					
	Nickel, Ni, 58.6					

Chemical Notation.

Chemical notation is the recording, by means of symbols, the chemical changes, or the migrations of atoms or molecules in either the formation or the resolution of compounds; in short, it is the symbolizing of chemical facts. We have already employed the first capital letter of the Latin or Greek names of the elements as shorthand for the whole name. Thus, for hydrogen or oxygen, we simply write H and O. This is not the only function of the symbol (to represent the name of the element); it means, furthermore, *one atom* of the element, and third, *one volume* of the element in the gaseous condition. A fourth office is to represent a *constant combining weight*—i.e., the atomic weight of the element for which it stands.

Figures are used to modify some of these functions. Large figures placed before a symbol multiply the atoms, and therefore also the atomic weights and the volumes in gaseous condition. Thus, 2H means two atoms of hydrogen, two volumes, and twice the atomic weight also. A small figure written to the right of the symbol, below, has the same significance; thus H_2 is the same as 2H . A figure preceding multiplies all the symbols that follow, as 3HCl means three of hydrogen and three of chlorine, but a small figure at the lower right of a symbol multiplies only the symbol to which it is attached. It would really be incorrect to write 2H , as we would thereby mean two atoms of hydrogen uncombined, which cannot thus exist; atoms, even of one kind, will combine to form molecules, two atoms usually forming one molecule, which molecule in the case of hydrogen or oxygen would be expressed: H_2 or O_2 . The third function of a symbol, that of representing one

volume of the element in the state of gas, will perhaps not be so easily understood by the student. By referring to the fourth function, representing the atomic weight, he may the better understand it. In learning the atomic weights, the student was informed that they were obtained by weighing *equal volumes* of the elements under like conditions of pressure, etc., after they had been caused to enter the gaseous condition, and that the number expressing the number of times the volume of any one gas was heavier than an equal volume of hydrogen, the standard, expressed the atomic weight of that element, for equal volumes of gases in the case of elementary bodies contain equal numbers of molecules; and because these, the molecules, contain two atoms the volumes must contain equal numbers of atoms. The extension or bulk of the volume is arbitrary, so long as it be equal to that with which it is compared.

The *position* of symbols, too, has some significance. H_2O means not only two atoms of hydrogen and one of oxygen, etc., but also that the H and O are united by chemical force and in the proportion expressed by the figures 2 and 1, respectively. (1 is always understood, and therefore never written.) Written $2H_2 + O_2$, it would indicate a mere mechanical mixture of two molecules of hydrogen and one molecule of oxygen, but placed closely together it is understood that they represent the compound resulting when hydrogen and oxygen unite chemically.

An *atom* is represented by a single symbol, but a *molecule*, which consists of not less than two atoms, is represented by as many symbols as it has kinds of atoms in its composition, and the collection of symbols representing the composition of a molecule is called its *formula*. The formula for a molecule of an *element* has only a single

symbol, representing the kind of atoms, together with the small figure at the lower right side denoting the number of atoms; thus, H_2 , O_2 , Cl_2 denote molecules of hydrogen, oxygen and chlorine, respectively. The molecules of some elements have more than two atoms; phosphorus, P, is believed to have four.

The formula for the molecule of a *compound* must have not less than two kinds of atoms. Calcium carbonate, common chalk, contains calcium, carbon and oxygen, and its formula is $CaCO_3$ —that is, one atom of calcium, one of carbon, and three of oxygen—and these elements are united in the proportions (definite proportions) expressed by their atomic weights. The formula for the molecule of water is H_2O , for hydrochloric acid, HCl . The latter may be made by inducing chemical union between the constituent elements; this union may be thus recorded: $H_2 + Cl_2 = 2HCl$, and such a series of formulas is called an *equation*. An equation is, therefore, the illustration by means of symbols of the chemical union or resolution of atoms or molecules.

We say *chemical resolution*, for it is just as possible to *resolve* a compound into its constituent elements as it is to cause elements to unite. $2HCl = H_2 + Cl_2$ would be an equation illustrating chemical resolution. In all equations the kinds and numbers of atoms to the left of the sign of equality must always equal those to the right of it.

It may be appropriately mentioned here that chemical *resolution* is also termed *analysis*, and chemical *union*, *synthesis*. Analysis and synthesis have usually given to them a broader significance. Analysis does not necessarily mean alone to break up compound substances into their elements, it means also that a complex combination may

be broken up into two or more simpler ones. Chemical processes and methods by which we recognize an element, or a group of elements, are termed chemical analyses, whereas chemical synthesis is the putting together, by means of chemical force, elements to form compounds, or uniting simpler compounds to form more complex ones.

By heating mercuric oxide, HgO , the two elements become dissociated. The analysis may be thus expressed: $2\text{HgO} + \text{heat} = \text{Hg}_2 + \text{O}_2$. This is one form of analysis; another may be illustrated by heating calcium carbonate: $\text{CaCO}_3 + \text{heat} = \text{CaO} + \text{CO}_2$. In the former the product of the analysis is the elements; in the latter, compounds which are simpler than the one they made up.

A simple form of synthesis is illustrated by bringing iodine and iron in contact for some time—iodide of iron is formed according to this equation: $\text{Fe}_2 + 2\text{I}_2 = 2\text{FeI}_2$. Another form of synthesis in which simple compounds are employed may be symbolized in this equation: $\text{SO}_2 + \text{H}_2\text{O} = \text{H}_2\text{SO}_3$. The sulphur dioxide unites with the hydrogen oxide, water, to form sulphurous acid.

Much more, for which the student is not yet ready, will be said about chemical notation and *chemical nomenclature* as the student advances.

The functions of symbols, formulas and equation may be briefly summarized:

A symbol is: (1) shorthand for the name of the element which it represents; (2) it represents one atom of that element; (3) one volume of that element in the gaseous condition; (4) it means a definite chemical proportion or atomic weight.

A formula gives (1) the names of the elements in the molecule; (2) the number of atoms in the molecule; (3)

the molecular weight or sum of the atomic weights; (4) it represents two volumes of the substance in state of gas, and (5) denotes that the atoms in the molecule are united by chemical force.

The function of an equation is to express by symbols the changes that take place among atoms and molecules in analysis or synthesis.

The Non-metallic Elements.

BEFORE entering upon the study of the non-metallic elements it is advisable for the student to review all he has studied up to the present time, so as to insure a thorough understanding of that which is to follow. It is imperative to have a good foundation to build upon, and it will be well worth the student's time to read over several times the principles of chemistry considered in previous pages.

The physical differences between non-metals and metals are obvious. Every one knows that iron and silver are metals. We can tell by their appearance, weight and lustre. It is just as obvious that the constituents of the air, oxygen and nitrogen, are not metals, and because they are *not metals* we call them *non-metals*. There is, however, no very clear line of distinction in Nature between these two subdivisions of the elements as they pass gradually into each other. The *metalloids*, bismuth, arsenic and antimony, may be classed in either division. They are called metalloids because they are *like* the metals in some respects.

Oxygen.

History.—It is appropriate that we begin the study of the chemical bodies with the most remarkable and important element—oxygen. It was discovered simultaneously in 1774 by Dr. Priestley, of England, and a young apothecary by the name of Scheele, in Sweden. Lavoisier, the distinguished French chemist, proved the identity of the gas discovered by Priestley and Scheele, but did not discover it, as is stated by some. Priestley was one day concentrating

the sun's rays by means of a burning glass, and directed the focus upon some red oxide of mercury, when, to his surprise, the oxide began to resolve itself into a gas and metallic mercury. The gas was afterward proven to be an element. Scheele, about the same time, heated a quantity of dioxide of manganese—a stone occurring abundantly in the locality of his home, and called *braunstein*, i.e., brown-stone—with a result similar to that which Priestley had obtained. It is remarkable that this gas should have been discovered by two different men, independent of one another, at the same time. The discovery of oxygen marked the beginning of the development of chemical science—it was, indeed, the birth of chemistry as a science. It was the beginning of the unveiling of Nature's most hidden secrets—the deliverance of chemistry from the bondage of the ridiculous, pretended science of alchemy.

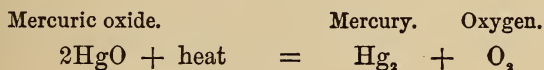
Occurrence in Nature.—Oxygen occurs everywhere in Nature. The three kingdoms are largely made up of it. It is the most widely-distributed and the most abundant element we have. The universe is computed to be made up of:

Oxygen.....	48 per cent.
Silicon.....	36 “ “
Aluminum.....	6 “ “
Calcium	} together.... 9 “ “
Magnesium	
Sodium	
Potassium	
Iron	
The rest of the elements.....	1 “ “
100 per cent.	

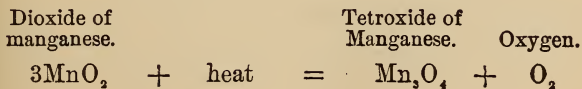
Water is 88 per cent. oxygen; air, 20 per cent. Fifty per cent. of all the rocks is oxygen; from 60 to 80 per cent. of plants and animals is oxygen. In short, it is present almost everywhere and in everything.

We know that oxygen is continually being produced by plants, these taking in the carbonic acid gas present in the air, resolving this into its constituent elements, carbon and oxygen, retaining the former and fixing it in their tissue, and liberating or setting free the oxygen.

Preparation.—It may be prepared to-day as Priestley and Scheele prepared it over a century ago. By heating mercuric oxide it is readily obtained, according to this equation:



It may be heated in a test-tube with a perforated cork inserted, through which a bent glass tube passes, to afford a means of collecting the gas as it is generated. To collect the gas it is necessary to pass it into an inverted vessel, as a bottle filled with water, the mouth of the bottle being under water. The gas will pass upward and displace the water—that is, it is collected over water. Heating dioxide of manganese will produce the same result, according to this equation:



A more ready method is heating chlorate of potassium, contained in a test-tube or other suitable apparatus with

the necessary fitting. This equation expresses the liberation of oxygen by this method:



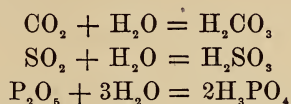
In the latter method less heat is required than in either of the former two, but care must be exercised because of the explosiveness of the chlorate.

If the chlorate is mixed with an equal weight of dioxide of manganese the evolution of oxygen occurs at a very low temperature, so that very little heat is required. In this mixture the dioxide of manganese does not yield any gas, as it decomposes at a much higher temperature; it acts by its presence only. Such substances, which take no part in a change like the above, are sometimes said to act by *catalysis*—that is, by their presence only. Their presence furnishes angles, and angles always facilitate the escape of gases.

Physical Properties.—Oxygen is a colorless, invisible gas, without odor, weighing 16 times as much as hydrogen, wherefore its atomic weight or specific gravity is 16. Its molecular weight is 32—twice its atomic weight. It has been liquefied by extreme pressure and reduction of temperature. It is soluble to the extent of 5 per cent. in water. This 5 per cent. sustains the life of fishes.

Chemical Properties.—Oxygen has such chemical affinity that it combines with every element with the exception of fluorine. It *oxidizes* all the elements. The act of combination is called *oxidation*, and the products formed, *oxides*. If carbon, sulphur or phosphorus be burned in oxygen, the resulting oxides, CO_2 , SO_2 and P_2O_5 , when dis-

solved in water form the corresponding acids, carbonic, sulphurous and phosphoric, thus:



This fact was noticed by the early experimenters, and believing they had found the "acid generator," they gave it the name "oxygen," which means "acid generator." Later chemists learned, though, that it is not oxygen but hydrogen (which means "water generator") which is the essential constituent of acids.

BURNING is a chemical process, oxidation, in which the chemical union between the burning body and oxygen is so intense that heat and flame are evolved. The more rapid and complete the oxidation the more luminous is the flame produced. Such a form of oxidation is termed *combustion*.

Substances which burn in air burn with more brilliancy in oxygen, which fact is employed in recognizing oxygen. A glowing piece of wood is instantly relighted into a bright flame in oxygen. Oxygen is the supporter of life; taken in through the lungs it causes the slow combustion of various organic substances in the living animal. It slowly oxidizes these bodies, which are mainly refuse, changing them into gases, which are eliminated during exhalation.

It is not necessary for the oxygen to be in its elementary condition to enter into combination or to oxidize bodies. Some substances contain the oxygen so loosely combined that they readily part with it when they come in contact with bodies capable of being oxidized.

The symbol for oxygen is O; atomic weight, 16; molecular weight, 32.

Uses and Compounds in Pharmacy and Medicine.—There are no direct uses for oxygen in pharmacy, excepting that it is kept in some pharmacies for physicians' use. Its compounds with other elements receive attention under the mineral acids and elsewhere.

Tests and Experiments.—When things burn in air they take up or unite with oxygen. Coal, when it burns, takes up or unites with the oxygen of the air, forming carbonic acid gas, thus: $C_2 + 2O_2 = 2CO_2$. This equation represents all the usual forms of combustion, and combustion is generally defined as the union of an element (usually carbon) with oxygen, with the production of a flame. The products of such combustion are *oxides*. It must be remembered that one-fifth of the air is oxygen, and that if combustion takes place with such an amount of oxygen, it will be very much more intense in pure oxygen. All substances which burn in air burn with very greatly increased brilliancy in the pure gas. If we should introduce a glowing piece of wood into oxygen, prepared as above, it would immediately relight and burn with extreme vividness. Charcoal, which ordinarily only glows, burns with a very bright flame in oxygen, forming CO_2 , carbonic acid gas; sulphur gives a beautiful blue flame, forming sulphur dioxide (gas), SO_2 ; phosphorus burns with a blinding brightness in oxygen, producing phosphoric oxide, P_2O_5 ; iron burns with beautiful scintillations, forming ferric oxide, Fe_2O_3 . Substances ordinarily incombustible in air burn with surprising brilliancy in oxygen. In all these cases the effects are due simply to the union of the burning body with the oxygen, forming oxides. This union is in obedience to the laws of chemical union already depicted, and if we would weigh the bodies and the oxygen in each case, we would

find these laws verified experimentally. These experiments are all *oxidations*, and the products formed *oxides*.

Ozone.

Ozone is a peculiar modification of oxygen. If a series of electric discharges is passed through oxygen the latter becomes modified in a manner that makes ozone of it. Ozone is an allotropic (another form) form of oxygen, a form which is sometimes designated as the active state of oxygen, while the ordinary oxygen is in the passive state.

Ozone has a characteristic odor, especially noticeable after electrical discharges. Oxygen becomes diminished in volume if a number of electric discharges are passed through it, and experiments have proved that ozone is one and a half times as heavy as oxygen, three volumes of oxygen becoming two volumes of ozone. The atomic weight of ozone is correspondingly found to be 24, oxygen being 16; or the molecule of ozone may be looked upon as

containing 3 atoms of oxygen $O_3 = \begin{array}{c} O \\ \diagup \quad \diagdown \\ O-O \end{array}$ while the molecule of oxygen contains only 2 atoms, $O_2 = O-O$.

Ozone has very marked oxidizing properties, exceeding those of oxygen. It occurs in the air to some extent and may be detected in it by its power to liberate iodine from potassium iodide by oxidizing the potassium. For this experiment white filtering-paper may be saturated with solution of potassium iodide containing a little boiled starch. The dry paper when it comes in contact with ozone becomes blue, owing to the liberation of iodine from the potassium iodide, and the reaction ensuing between the iodine and starch, which produce the blue iodide of starch. Other

gases exercise this same oxidizing effect (e.g., the oxides of nitrogen, and hydrogen dioxide). Ozone may be distinguished from these in that it is converted into ordinary oxygen when passed through a heated tube, which is not the case with the other oxidizing gases. Ozone is also produced when strong sulphuric acid acts upon potassium permanganate, or when phosphorus is slowly oxidized in moist air.

Hydrogen.

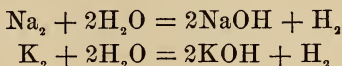
History.—Hydrogen was discovered by Cavendish in 1766, who also was first to describe it as an element and to prove that it was a constituent of water. By some it is stated that Paracelsus obtained it in the sixteenth century, but Cavendish first recognized its elementary nature.

Occurrence in Nature.—Hydrogen is never found free, on account of its intense chemical affinity, but in countless combinations it exists widely distributed in Nature. It forms one-ninth by weight of water; with nitrogen it forms ammonia, NH_3 , which is of considerable pharmaceutical importance; it constitutes a large proportion of all organized bodies, and it is found in products of decomposition, in shooting stars and in volcanic gases in the free state, but does not remain so long, usually combining with the oxygen in the air.

Source.—The usual source of hydrogen is water, from which it is obtained in various ways, some of which are described in the following paragraphs.

Preparation.—If an electric current is passed through water it effects a decomposition of the water according to this equation: $2\text{H}_2\text{O} + \text{electricity} = 2\text{H}_2 + \text{O}_2$.

Sodium and potassium as elements also effect the decomposition of water, with liberation of hydrogen and formation of hydroxides of the respective bases, thus:



A third and the most convenient method of preparing hydrogen is by means of zinc and hydrochloric or sulphuric acid: $\text{Zn}_2 + 2\text{H}_2\text{SO}_4 = 2\text{ZnSO}_4 + 2\text{H}_2$.

The zinc, preferably in form of granules, is introduced into a small flask into which is fitted a rubber stopper, through which passes a glass tube, bent at right angles just as it leaves the stopper. The tube should terminate in the flask just below the lower end of the stopper, and should be drawn to a point at its other end. When the sulphuric acid (diluted) comes in contact with the zinc the evolution of hydrogen begins, the gas finding its exit through the tube. It may, like oxygen, be collected over water. The zinc in the flask becomes dissolved, forming sulphate of zinc.

There are many other methods for preparing hydrogen, but this is the one usually employed for experimental purposes.

Physical Properties.—Hydrogen is a colorless, odorless, invisible gas, *very inflammable*, and slightly soluble in water. It is the lightest of all known bodies, and is therefore employed as the standard of comparison for the atomic weights of all the elements. It is an exceedingly diffusive gas, more so than any other; it will escape through joints of apparatus which are impervious to other gases. It is well to remember *that all gases are diffusive, and that they diffuse in inverse proportion to the square root of their den-*

sity or specific gravity. Thus oxygen, having a specific gravity or density of 16, will diffuse but one-fourth as fast as hydrogen:

O = 16, of which the square root is 4;

H = 1, of which the square root is 1;

or, in other words, hydrogen diffuses four times as fast as oxygen. On account of its extreme lightness it is sometimes used for inflating balloons. Hydrogen is a non-supporter of combustion, but combustible itself. It will not, like oxygen, support life.

Chemical Properties—Tests and Experiments.—If the hydrogen, prepared as above from zinc and sulphuric acid and passed through the glass tube, be brought in contact with an ignited match it will burn with a yellowish, nearly non-luminous, but very hot flame. (Care should be taken that all the air has been expelled from the flask to prevent explosion.) This flame is due to the chemical combination of hydrogen with the oxygen in the air, forming water in this wise: $2\text{H}_2 + \text{O}_2 = 2\text{H}_2\text{O}$. When hydrogen burns it always produces water. There is a very strong affinity between these two gases, and if mixed in any quantity and ignited, violent explosion taken place. In the above experiment only a small portion of hydrogen comes in contact at a time with oxygen. The *oxy-hydrogen* blowpipe is an arrangement whereby two volumes of hydrogen are made to meet one volume of oxygen at a point which is the terminus of two tubes conveying the two gases in proper proportions to form water. These proportions are the most economical, and produce greatest heat and light; but they may vary within wide limits. The union of the two gases produces a very high tem-

perature, one hot enough to melt iron or almost anything. This is a vivid illustration of the intensity with which chemical combination sometimes takes place.

If soap-bubbles be blown from a bag containing a mixture of 2 parts hydrogen and 1 part oxygen, and touched with a match or a candle, they explode with a detonation like that of a pistol-shot. Chemical union is not always spontaneous; in this, as in some other cases, it is necessary to employ an inducing-agent, the flame in this case.

If a jar be filled with hydrogen and held with the mouth downward the hydrogen will not escape very readily. A lighted candle introduced into the jar will become extinguished, while the hydrogen will burn at the mouth of the jar. The candle may be withdrawn, when it will relight, becoming extinguished a second time if introduced into the jar again. This shows hydrogen to be *inflammable but not a supporter of combustion*.

A great amount of chemical energy is possessed by hydrogen; one part of it by weight will neutralize 35.5 of chlorine, 80 of bromine and 126 of iodine, *or as much of any of the elements by weight as is represented by their atomic weights divided by their quantivalence*. It will extinguish, therefore, 16 of sulphur—the atomic weight of sulphur being 32, this divided by its quantivalence, 2, in this case, giving 16.

Hydrogen is the base in all acids, and may be replaced in them by the alkalies or metals to form salts. In hydrochloric acid, HCl , the hydrogen may be replaced by sodium, Na , to form sodium chloride, NaCl . Potassium bromide, KBr , as a further illustration, is hydrobromic acid, HBr , in which the H is replaced by potassium, K .

IS HYDROGEN A METAL?—We are treating of hydrogen now as among the non-metals. It has long been a question whether it is a metal or a non-metal. Many of its chemical properties would class it with the metals, but its physical properties would place it among the non-metals. Some chemists suggest that it is a metal which exists at the prevailing temperature in the form of a gas, and that if the temperature could be reduced far enough it might be obtained in the form of a liquid or solid. In view of this and the analogy of its behavior to the other metals, many accept hydrogen as a metal. It is here studied among the non-metals for convenience.

In the beginning of this course the student learned that all bodies exist in one of the three forms—gaseous, liquid or solid—and that the condition depends upon the predominance of one or the other of the molecular forces, cohesion and repulsion. It is appropriate to mention here that these molecular forces are in their turn dependent upon the temperature, heat destroying or lessening the cohesive and increasing the repulsive force, cold increasing the cohesive and lessening the expansive force.

Different bodies are differently affected by different temperatures, mercury, for example, being liquid at 32° F., while water becomes solid at that temperature.

Uses in Pharmacy.—There are no specific uses in pharmacy for hydrogen, excepting perhaps its employment in some chemical processes, as in the preparation of reduced iron from ferric hydroxide, etc., or in the application of tests for the purity of some of the pharmacopœial products.

Compounds.—Of the compounds of hydrogen water

ranks first and the acids next, the carbohydrates and hydrocarbons (unions with carbon), in the organic world, following. It forms *hydrides* with some of the elements— PH_3 , AsH_3 , SbH_3 —though it is contended by some that the term is not correctly expressive of the form of combination. It is preferred to call these combinations, PH_3 , phosphoretted hydrogen; AsH_3 , arseniuretted hydrogen, etc. With oxygen it forms, besides water, another compound, H_2O_2 , the hydrogen dioxide or peroxide of hydrogen. The prefix *per* means *high, above, to the fullest extent*, and hence a *peroxide* is a combination containing more oxygen than another compound containing the same elements: H_2O , water, is an oxide of hydrogen; H_2O_2 , containing more oxygen, is the peroxide. The same applies to other elements besides oxygen—there are perchlorides, persulphates, etc.

Hydroxides—often called *hydrates*—are also compounds containing hydrogen; they are usually looked upon as water, H_2O , $\text{H}-\text{O}-\text{H}$, in which one of the hydrogen atoms is replaced by some other element or base. Potassium hydroxide would be $\text{K}-\text{O}-\text{H}$, the K taking the place of one of the hydrogen atoms.

Water.

Water being an oxide of hydrogen, it is appropriate to treat of it under hydrogen.

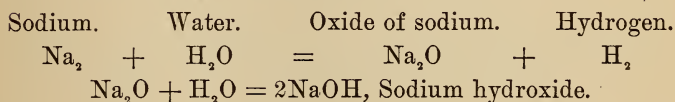
The importance of water in the economy of nature must be apparent to all, in that it is one of the essentials in the maintenance of all living things—without it there could be no life. It is the most abundant substance in nature

and in its changes of form from the solid ice to its invisible vapor involves the very history and destiny of the earth. Water is concerned in the majority of chemical processes, and is as indispensable in the laboratory of the chemist as in that of Nature. The indispensability of water was apparent to the ancients, and we readily excuse them for having classified water with their elements, which we all know to have been air, fire, earth and water. The elements of the ancients were supposed to be transformable one to the other; it was believed that water could be changed to earth, and *vice versa*, until a chemist disproved that belief by distilling and redistilling a weighed quantity of water in air-tight stills and condensers for more than 100 days, at the end of which time he found that the water had not lost weight nor volume. He employed distilled water, which had not been done by others, they always having obtained a residue which was composed of the fixed matter present in all waters, but which they supposed to be transformed water.

Composition.—Water is composed of 8 parts by weight of oxygen and 1 part of hydrogen, or, by volume, of 2 of hydrogen and 1 of oxygen; its composition is expressed by the formula H_2O . That this is so may be proven in several ways, the simplest being the decomposition of water by the electric current. The water is contained in an apparatus so arranged that the two gases are set free in separate graduated tubes, which are filled with water. The volume of each gas may be read off, after decomposition, and will be found to be twice as great for the hydrogen as for the oxygen. If the gases, in the exact proportion as given off, are mixed and ignited they will combine with a loud explosion, reforming pure water.

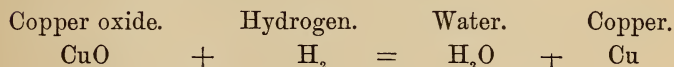
We have thus two methods of proving the composition of water—the one by analysis, the other by synthesis, or “putting together.”

If metallic sodium or potassium be thrown upon water the latter will decompose with such violence as to produce a flame, usually. The oxygen which is liberated seizes upon the metal and forms its oxide, the oxide uniting with water to form the hydroxide:



Iron, or zinc, when heated to redness, has the same action on water.

A process upon which much reliance can be placed for demonstrating the composition of water is that in which pure oxide of copper is reduced by hydrogen, and the water so formed is collected and weighed.



By noting the weight of the oxide of copper before and of the copper after the operation, weighing the hydrogen and finally the water, we can tell the composition by weight of the water formed.

The elements of water are separated and reunited in numberless operations in chemistry, and the same thing is going on continually in plants and animals. Indeed, water, constituting four-fifths of the vegetable kingdom and three-fourths of the animal, is the prime condition of all organization, and the maintenance and continuation of or-

ganic life are largely dependent upon its many transformations and decompositions.

General Properties.—Water is a colorless, transparent, tasteless, inodorous, volatile liquid, or solid if the temperature is below 32° F. It is practically incompressible. It evaporates spontaneously at ordinary temperatures, boils at 212° F. (100° C.), and freezes at 32° F. It does not expand or contract equally, having its maximum density at 39.2° F. (4° C.), and expanding 10 per cent in becoming ice. The weight of 1 c.c. at that temperature is 1 gramme, which is the standard of weight in the metric system. A cubic inch of water weighs at ordinary temperature 252.45 grains, whereas a cubic foot weighs about 1000 avoirdupois ounces. Water in small bulk is colorless, but in mass it is blue, like the atmosphere. The admixture of the slightest quantity of some substances destroys the color.

Water never occurs in nature in a state of absolute purity. It always contains some dissolved matter, usually saline, sometimes organic.

All natural waters contain in solution the bicarbonates, chlorides and sulphates of sodium, potassium, calcium, magnesium, and often iron and silica. The only absolutely pure water is that obtained by distillation. Rain-water, even, evaporated from the sea and condensed in the air, descending through the atmosphere, becomes charged with the soluble impurities present in the air, such as ammonia, carbonic acid gas, etc., and hence is not absolutely pure.

Chemical Properties.—Water forms *hydroxides* when brought in contact with many of the oxides, the combinations being in some instances so intense that considerable

heat is evolved. The hydroxides contain the hydrogen and oxygen in the proportion to form water, but they do not exist as water, the atoms suffering a change of arrangement during the union with the bases. Thus, caustic potassa, KOH, or potassic hydroxide, may be looked upon as being composed of $K_2O + H_2O (= 2KOH)$.

Sodium hydroxide, $2NaOH$, is Na_2O and H_2O

Calcium hydroxide, $Ca(OH)_2$, is CaO and H_2O

The water is retained in these compounds with such tenacity, that a high heat is necessary to remove it. In case of calcium hydroxide it requires a red heat to expel the water and convert it into the oxide again, while in many other instances the water is held with such force that it cannot be expelled by heat at all.

The term *hydrate* is often used for *hydroxide*, but it should not be, as *hydrate* has another specific meaning. Hydrates are bodies, salts, for example, which unite with water without causing an alteration of atomic arrangement. The water in such combinations is termed *water of crystallization*, and may be easily removed by comparatively low heat. Sodium carbonate, for example, has ten molecules of water, which latter may be said to be in juxtaposition to the sodium carbonate, and which can be separated without altering the composition of carbonate proper. The formula is $Na_2CO_3 \cdot 10H_2O$. Heat will remove the $10H_2O$, leaving the Na_2CO_3 .

Salts which do not contain any water of crystallization are *anhydrous*, so common salt, sodium chloride, $NaCl$; potassium chlorate, $KClO_3$, etc., are *anhydrous*. *Cryohydrates* are hydrates which exist only at temperatures below

the freezing point. Sodium chloride, which at ordinary temperatures is anhydrous, becomes a cryohydrate when solidified at about 40 degrees below the freezing point, when it has this formula, $2\text{NaCl} \cdot 21\text{H}_2\text{O}$.

When the water of crystallization is so feebly united with the salt that it gradually separates at the ordinary temperature, the salt is said to be *efflorescent*, and the change is called *efflorescence*. Sodium carbonate crumbles because it is efflorescent—that is, the water is instrumental in the formation of crystals, which lose their shape when the water is expelled, or when it separates spontaneously.

On the other hand, some bodies take up water from the air and become liquefied; such bodies are said to be *deliquescent*, and the change is called *deliquescence*. Calcium chloride is a good example of a deliquescent salt.

Water is the very best solvent we have. As a rule, salts are more soluble in hot water than in cold; notable exceptions are common salt, which is equally soluble in water at any temperature, and many calcium salts, which are less soluble in hot than in cold water. Water dissolves many of the gases, and therefore also the air, to some extent, but the air dissolved in water has not the same composition as that constituting the atmosphere, which is one-fifth oxygen and four-fifths nitrogen. This is because the oxygen is much more soluble in water than the nitrogen.

Natural Waters.

These may for convenience be divided into two classes, the atmospheric and the terrestrial, and these classes still further subdivided, thus:

Atmospheric.	Terrestrial.
Rain	Ocean
Snow	Seas
Hail	Lakes
Fog	Lakes
Dew	Rivers
Watery vapor	Creeks
	Brooks
	Springs
	Wells

} Salt.

} Fresh.

Atmospheric Waters.—Of the atmospheric waters we need not say much. They are all the result of the condensation of the water evaporated from the surface of the earth.

The great specific heat of water has a powerful controlling influence on the climate. Perhaps the student does not know what *specific heat* means, and a brief exposition of it may be in order. Equal weights of different bodies at the same temperature require different amounts of heat to raise them to a given degree of temperature. A certain amount of heat is required, for example, to raise 1 pound of water at 100° to 101°; that amount is not the same for all bodies, but varies for every body. The amount of heat required to raise a body one degree in temperature, compared to the amount of heat required to raise an equal weight of water through 1°, is that body's *specific heat*. There are several methods of obtaining the specific heats of bodies. If 1 pound of water having a temperature of

100° be mixed with 1 pound of water of 40°, the 2 pounds will have a mean temperature of 70° : $100 + 40 = 140$, $\div 2 = 70$. The first pound of water gives 30° to the water of 40°, making that and becoming itself 70°. This is true if equal portions of any one liquid of different temperatures are mixed, but if unlike bodies having different temperatures are mixed, a mean temperature does not result.

If 1 pound of water of 100° is shaken with 1 pound of mercury at 40°, then, these being unlike bodies, instead of a mean temperate of 70°, the temperature will become 98°. The water lost 2°, an amount sufficient to raise the temperature of an equal weight of mercury 58°; that is, from 40° to 98°. It is evident from this that the specific heat of mercury is only $\frac{2}{58}$, if water is 1. In the first case the 30° of heat which the 100°-water lost raised the temperature of water of 40° to 70°, or every 1° of heat from the water of 100° raised the water at 40° 1°, but in the second case, since the 2° which the water lost were sufficient to raise the pound of mercury through 58°, to raise mercury through 1° would require $\frac{2}{58}$ as much heat as water requires.

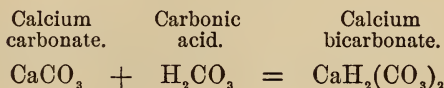
Water has the highest specific heat of any body known. The stability of Nature is dependent upon this high specific heat of water. Were it not so, the rapid changes in temperature which would ensue would unfit this globe for habitation. If water could acquire heat with as much facility as mercury does, the seas and water everywhere would freeze and thaw so rapidly and often that all living things would perish from want of adaptation.

Terrestrial Waters include all that are on or in the earth; they are divided into the *salt* and *fresh* waters.

The ocean, inland seas and some lakes are salt, while all the other divisions are fresh; that is, without salt. The *fresh* water is further classified into *soft* and *hard* water.

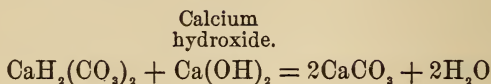
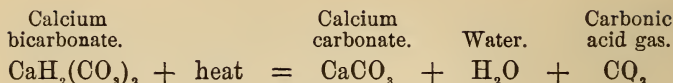
Hard Waters.—As we have learned, all natural waters contain dissolved matter; this is especially true of terrestrial water, which, while traversing the surface or interior of the earth, becomes charged with the soluble portions of whatever it comes in contact with. If the water contains carbonic acid gas, CO_2 , its property of dissolving certain bodies is increased. All rocks are soluble to a slight extent in water, but those containing lime, magnesia or iron are more soluble in water containing carbonic acid gas.

A water which happens to contain carbonic acid gas freely dissolves limestone, which is *calcium carbonate*, by converting it into the *soluble bicarbonate*:

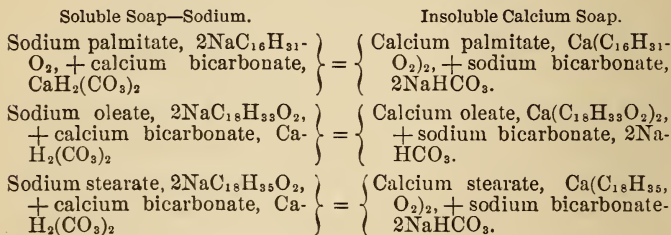


Such water would be called a *hard* water. Water also dissolves out of the rocks and soil the *sulphate of calcium*, CaSO_4 , and thereby becomes another variety of hard water, known as *permanently* hard, while the former, containing the bicarbonate of calcium, is termed *temporarily* hard. These terms express the nature of the water relative to the kind of calcium salt present. Because we can remove the hardness of the water containing the bicarbonate in solution, it is called the *temporary* hardness, while if due to the sulphate it is *permanent*, because it cannot be remedied.

The temporary hardness of water may be removed by boiling or by adding lime-water, a solution of $\text{Ca}(\text{OH})_2$, as shown by the following equations:



The soluble *bicarbonate* becomes the insoluble carbonate in both instances, and this precipitates, and in so doing renders the water soft, for the hardness is caused only by the calcium salt being *in solution*. The use of soap with such water is attended with great loss of soap in the chemical change which takes place in the endeavor to form a lather. Soap with soft water gives a lather at once, but with hard water only after all the calcium salt has been converted into the insoluble calcium soap. The following equations will illustrate the change:



Ordinary soap is a chemical compound consisting of three definite salts—the palmitate, oleate and stearate of sodium, which, with the soluble lime salt, become converted into the corresponding calcium salts, which are insoluble. The greasy scum which floats on hard water which has been treated with soap, is the light and insoluble calcium soap, which gives to the touch the sensation of hardness, whence the name.

Mineral Waters.—We have learned, and it is worth repeating, that all natural waters contain dissolved matter, and that spring-water, especially, always contains the chlorides, sulphates and bicarbonates of sodium, potassium, calcium, magnesium and sometimes iron and silica. A *mineral water* is one which contains one of these usual constituents in an abnormal or unusual quantity or a substance which is not usually found in natural waters. All the mineral waters scattered over the earth are simply spring waters impregnated, to a greater or less extent, with soluble, substances to which medicinal virtues are attributed. Some waters hold carbonic acid gas in solution, when they are termed *sparkling* or *effervescent*, while those without that gas are termed *still*. When iron is present the water is known as *ferruginous* or *chalybeate*; others are *alkaline*, while some are *acid*. *Borax waters* are so called because they contain borax, while the name *alum* water is due to the presence of alum in such waters. When much sulphate of sodium (Glauber's salt) is present, or sulphate of magnesium or other *salt*, we term the water *saline*. *Sulphur waters* are so called because they contain sulphur, usually in the form of soluble sulphides. Others contain notable quantities of iodine, some bromine, others arsenic, still others lithia, etc.

Most of these mineral waters must be used as they are taken from the springs, but a number of them may be evaporated to a small bulk.

Well-water.—Wells, unless situated distant from habitations, may become dangerous. The soil in which they are dug or driven is usually pervious, so that they may become contaminated with anything with which water trickling from the surface through the soil into them may come in

contact. Decaying organic matter, the refuse from stables, etc., are dissolved by the water falling upon them as rain, which in its descent through the soil will find its way into the well. This organic matter does not necessarily unfit the water for domestic purposes; the danger lies in its affording a breeding medium for disease germs, which, it is well known, develop rapidly when brought into the system.

To test well water or other drinking water for its purity comes within the domain of the pharmacist, though a complete and exhaustive analysis should be referred to a skilled chemist. The substances to be looked for primarily are organic matter, albuminoid matter, ammonia and nitrates and nitrites. The nitrogen compounds are usually more abundant if animal matter is present; they, of all other constituents, render water the most unwholesome. The following tests are given here, but the student may omit their application until he will have advanced a little farther.

1. *For Organic Matter.*—Color a portion of the sample distinctly with a solution of permanganate of potassium and add two or three drops of dilute sulphuric acid. If much organic matter is present the color of the permanganate becomes discharged almost immediately; if less, it takes a little longer to decolorize. If the color has not changed in 20 to 30 minutes it is safe to assume that organic matter is not present. This is a tolerably reliable test.

2. *For Nitrites.*—A little sulphuric acid added to the water containing nitrites forms nitrous acid, which is easily detected by its power to liberate iodine from a solution of iodide of potassium. A little starch paste, made by boiling starch with water, is mixed with a solution of iodide of

potassium and the whole added to the suspected sample, and if nitrites are present the starch will turn blue with the liberated iodine.

This indirect method is a means for detecting nitrites, if present in not too small a quantity.

3. *Nitrates* are detected by converting into nitric acid, which turns morphine red. A portion of the water is evaporated to dryness, the residue treated with a drop of sulphuric acid, which makes nitric acid of the nitrate, and a portion of morphine added. If nitrates are present a red color will ensue.

4. *For Ammonia*.—Nessler's reagent affords by far the best test. The solution may be made by dissolving 18 grains of iodide of potassium in a little water, adding solution of mercuric chloride until the red iodide first formed redissolves upon agitation. To this are added 50 grains of caustic potassa, and distilled water to make 8 ounces. This reagent will detect 0.00375 grain of ammonia in a pint of water by giving yellow to reddish color.

Albuminoid Ammonia requires a more elaborate process for its detection. If all the above were found it is hardly necessary to go to the trouble of looking for albuminoids; the water would be unwholesome if they were not present. If it is required to test for them, nevertheless, Chapman and Wanklyn's test may be employed. If the water was found to contain ammonia, this must first be removed, as must also any urea that may be present. This is best done by distilling the water until it gives no further reaction with Nessler's solution. Then add a strong solution of caustic potassa and potassium permanganate, and examine again for ammonia. This test depends upon the fact that caustic potassa and permanganate of potassium cause animal mat-

ter, while still in an ambuminoid condition, to unite with hydrogen to form ammonia.

These tests are easily applied, but they are only tolerably reliable. If all respond readily, it is quite safe to assume that impurities are present. If, however, they do not respond satisfactorily, and an important issue is in question, the suspected sample should be referred to an analytical chemist.

The Purification of Waters.—Of the methods for purifying waters we have already become acquainted with those for hard water. The best general method is that of distillation; all fixed matter remains in the retort, while foreign volatile bodies may be dispelled by boiling previous to distillation. To render water perfectly pure it should be redistilled in silver vessels.

Mechanically suspended organic impurities in water are noxious, rendering water unwholesome and unfit for domestic or culinary purposes. It should be stated, however, that when these organic impurities are present (they may be from the drainage of houses, stables, from decaying leaves and animals, from the dust in the air, etc.) they are usually attended by a more or less efficient purifying agent in the form of animalculæ (or little animals), which feed upon them. These minute organisms may be viewed through a powerful microscope, when they will exhibit definite yet hideous shapes. There are about 30 species (or more) of them, all performing invaluable service in purifying water by consuming dead organic matter, reducing it to its ultimate harmless constituents—carbonic acid gas, ammonia and water. They themselves are harmless, yet, where they can exist, the pathogenic or disease-generating microbes may also flourish.

Boiling is a simple means of purifying water; it kills animal and vegetable matter (though not all disease germs), and expels gases. If lime is in solution it becomes precipitated.

Turbid or muddy water may be clarified by the addition of a little alum (ten grains to a gallon), which forms with bicarbonate of calcium, which is usually present, a hydroxide of aluminum, which in its descent carries the suspended matter with it mechanically.

Filtering through charcoal, sand, felt or brick, or through some other porous medium, also deprives water of suspended matter.

If a doubtful sample of water must be used for culinary purposes it should be boiled for half an hour and filtered; any contaminating matter will thereby probably become destroyed.

Peroxide of Hydrogen.

We leave now the monoxide of hydrogen, or water, to briefly consider the dioxide or peroxide of hydrogen, and with that will dismiss the compounds of hydrogen.

Peroxide of hydrogen is, as the name indicates, a higher oxide than water, having the formula H_2O_2 ; by some it is claimed to be a hydrate of oxygen, OH_2O .

It is a transparent liquid, colorless, of a specific gravity of 1.006 to 1.012. It volatilizes at ordinary temperatures, and when brought in contact with some bodies decomposes with violence. It is odorless, has an acidulous taste, and possesses pronounced bleaching properties. Because it is very unstable, the commercial product has added to it a small proportion of a mineral acid, which renders it somewhat more permanent.

It may be made from the peroxide of barium, BaO_2 , by treating with hydrochloric acid: thus, $\text{BaO}_2 + 2\text{HCl} = \text{BaCl}_2 + \text{H}_2\text{O}_2$. The BaCl_2 is precipitated with sulphate of silver, which produces the insoluble salts, silver chloride and barium sulphate, which are removed by decantation.

The aqueous solutions of the market are known as the 10 volume and 15-volume products. The 10-volume article, which is official in the U. S. Pharmacopœia, is known by that name because it contains 3 per cent by weight of the peroxide, which yields ten volumes of active oxygen. The oxygen is the active constituent doing the bleaching for which the solution is mostly used. The quantity of oxygen may be estimated by noting the quantity of acidulated permanganate of potassium which it decolorizes.

A variety known as the C. P. solution is used in medicine and surgery as a powerful antiseptic. Externally it is used as an application to ulcers, and as a dressing to all kinds of wounds. Its value, of course, depends upon the oxygen which it gives off.

Nitrogen.

History.—Nitrogen was not discovered until 1772, when Rutherford isolated a gas which he claimed to be an element, and which claim was subsequently proved to be true. At first it had two names given it, one, *nitrogen*, because of its source from nitre; the other, *azote*, because it was supposed to be poisonous, that word meaning without life. Animals kept in nitrogen for only a brief time would perish. We know now that death was not the result of the toxic properties of the nitrogen, but the result of the absence of oxygen, which latter is the life-sustaining principle of the air.

Occurrence in Nature.—Nitrogen is widely distributed in nature, being most abundant in a free condition in the atmosphere, of which it constitutes about four-fifths. It plays an important part in the vegetable kingdom, entering into many compounds and forming an essential constituent in others. Alkaloids contain it, as do ammonia and nitric acid and all cyanides. It is found, combined with potassium, as Calcutta nitre in India, as an efflorescence on the earth's surface. The Chili nitre, or nitrate of sodium, occurs in a like manner on the soil in Chili. Some varieties of coal also contain nitrogen.

Functions in Nature.—The chief function of nitrogen in the air is to dilute the more energetic oxygen. It is also essential to plants, which imbibe it in the form of ammonia.

Properties.—Nitrogen is a colorless, inodorous gas, without taste, and from a chemical point of view a very indifferent body—to combine it with other bodies we must employ a very circuitous method. It is incombustible and does not support combustion; a lighted match or taper introduced into it is immediately extinguished. It has been liquefied by Cailletet and Pictet and Dewar; its solubility in water is only very slight. It has at times 1 bond, at others 3, 5 and even 7; its atomic weight is 14; molecular weight, 28; symbol, N.

Preparation.—It may be prepared in a number of ways, but most commonly it is obtained from the atmosphere by withdrawing or using up the oxygen in it, leaving the nitrogen by itself.

A small piece of phosphorus floating in water on a small porcelain cup is ignited and a jar set over it immediately. The phosphorus in burning unites with the oxygen of the air, forming phosphoric oxide, P_2O_5 . The latter is very

soluble in water and is soon absorbed, leaving comparatively pure nitrogen in the jar.

Compounds.—With oxygen, nitrogen forms five compounds:

N_2O , nitrogen monoxide or hyponitrous oxide.

N_2O_2 , nitrogen dioxide.

N_2O_3 , nitrogen trioxide or nitrous oxide.

N_2O_4 , nitrogen quadroxide or tetroxide.

N_2O_5 , nitrogen pentoxide or nitric oxide.

Of these, three are capable of uniting with water to form acids:

$N_2O + H_2O = 2HNO$, hyponitrous acid.

$N_2O_3 + H_2O = 2HNO_2$, nitrous acid.

$N_2O_5 + H_2O = 2HNO_3$, nitric acid.

The N_2O , sometimes called laughing-gas (also, erroneously nitrous oxide), is easily prepared by heating ammonium nitrate: $NH_4NO_3 + \text{heat} = 2H_2O + N_2O$. This is a gas of a sweetish taste, colorless and almost inodorous, having the characteristic property of producing anæsthetic effects upon the system, for which it is employed in dentistry.

The N_2O_2 is produced when nitric acid is heated to decomposition or when it acts upon metals, or whenever it acts as an oxidizing agent. Nitric acid is a most powerful oxidizer; it always resolves itself into water, N_2O_2 and oxygen in this wise: $2HNO_3 + \text{heat} = N_2O_2 + 3O + H_2O$. The oxygen does the oxidizing in combining with other bodies. N_2O_2 is colorless, but when it comes in contact with the oxygen of the air it unites with it to form the tetroxide: $N_2O_2 + O_2 = N_2O_4$, which is of a reddish color.

The N_2O_5 with water forms nitric acid, as shown above. Two molecules of N_2O_2 are sometimes looked upon as being

composed of one molecule of N_2O and one of N_2O_3 : $N_2O + N_2O_3 = 2N_2O_2$. A similar view is taken by some of the composition of the N_2O_4 . The N_2O_3 and N_2O_5 together equal two molecules of N_2O_4 .

With hydrogen, nitrogen forms ammonia, NH_3 , which is constantly generated in nature during the decomposition of organic matter. The ammonia is also produced by the destructive distillation of organic matter, illustrated in the manufacture of illuminating gas, which is made by the destructive distillation of coal. Ammonia and its salts, as well as the acids of nitrogen, will be considered under other headings.

All alkaloids contain nitrogen, as do also cyanogen (CN) and its compounds. These will receive attention in their appropriate places. Many salts containing nitrogen are poisonous; notable exceptions are the nitrates of sodium and potassium.

Test for Nitrates.—If to a portion of strong sulphuric acid contained in a test-tube a freshly prepared solution of ferrous sulphate is gradually added so that it forms a layer above the acid, and a portion of a solution of a nitrate is slowly added and the tube gently agitated, a brown zone will form between the acid and the iron solution.

Oxygen and Nitrogen as the Atmosphere.

The invisible elastic envelope which surrounds the globe as a mantle is the *atmosphere*, the gaseous matter of which it is constituted being usually distinguished by the name of *air*. The air possesses no color or taste, nor can it be felt or grasped with the hand, yet, notwithstanding, it is material. Every one has noticed that when a funnel fits very tightly into the neck of a bottle, water poured into

it will not run through it into the bottle, the *air* within the bottle finding no outlet, and being material, according to the general property of matter termed impenetrability (in virtue of which no two bodies can occupy the same space at the same time), prevents the water from entering the bottle. When the funnel is slightly raised the water rushes into the bottle, the air being displaced at the same time and rushing out of the bottle.

Air has weight like all other forms of matter; a globe when weighed has a greater gravity than when the air is exhausted from it. To exhaust a vessel, a globe, e.g., of air, an instrument called an *air pump* is employed, which removes the air more or less thoroughly, causing a *vacuum*, though it is almost impossible to create an absolute vacuum. The difference in weight of the flask or globe before and after exhausting the air, the volume of its capacity taken into consideration, will show the weight of the air which the flask or globe will hold. It is thus ascertained that a cubic foot of air weighs about 538 grains; a room 40 feet square and 18 feet high contains about a ton of air. Air compared with hydrogen has a specific gravity of 14.4; water is about 825 times heavier than air. It is supposed that the air extends about 50 miles upward, although some claim only 10 miles, and others even 200 miles; more recent experiments favor the first-mentioned supposition. It is obvious that this tremendous ocean of air, at the bottom of which we live, must exert a vast pressure upon the surface of the earth.

This pressure amounts to 15 pounds upon every square inch of the earth's surface; it is not only directed downward, but in all directions—upward, sideward, etc. The weight of the air is not the same at all times, as it be-

comes condensed or rarified by atmospheric changes induced by winds, storms, etc.; it may be conceived to be moved by tides which sweep over the earth, and with these movements the pressure or weight changes, though never very much. The variations in the pressure may be measured, and the instrument by which we can record these changes is called a *barometer*. To make this instrument a glass tube about 33 or 34 inches in length, and closed at one end, is filled with mercury, the open end closed with one finger and inverted into a vessel containing mercury. The finger should not be released until the open end is fully under the mercury in the vessel. The mercury in the tube will sink until the column is about 30 inches from the surface of the liquid in the vessel. The weight of the air pressing upon the mercury in the vessel supports the mercury column at the given height—30 inches. One cubic inch of mercury weighs $\frac{1}{2}$ -pound, and if the tube were 1 inch square the 30 inches in height would weigh 15 pounds.

It is thus that it was ascertained that the pressure of the air is 15 pounds upon every square inch, for it requires a pressure of 15 pounds to support a column of mercury weighing 15 pounds. The column of mercury will rise or fall according to the pressure of the air; when the air is rare the barometer is low, when the air is condensed and heavy the barometer registers higher. This atmospheric pressure is often spoken of simply as “atmosphere;” two “atmospheres” would indicate a pressure of 30 pounds upon a square inch, three “atmospheres” one of 45 pounds, etc. The pressure decreases as the height increases, and increases as the distance downward from the surface of the earth increases. That is because as we ascend, the bulk of air above us becomes less, and as we descend, the bulk, and consequently also the weight, becomes greater.

The Chemical Constituents of the Air.

The atmosphere is a mixture of a number of gases, among which nitrogen and oxygen are the most abundant. Its composition by volume is computed to be:

Oxygen.....	20.81	Nitric acid.	} Traces.
Nitrogen.....	77.95	Ammonia.	
Carbonic acid gas.....	0.04	Sulphuretted hydrogen.	
Aqueous vapor.....	1.20	Carburetted hydrogen.	
	100.00	Sulphurous acid.	
		Organic vapors.	

The oxygen is the animal-life-sustaining principle of the air, while the nitrogen dilutes it so that animals can thrive in it—in pure oxygen we could not live. The oxygen oxidizes the refuse matter collected by the blood, in the lungs, converting much of it into carbonic acid gas, which is exhaled, and therefore is always present in the atmosphere. The composition of the air going into and coming out of the lungs is this:

<i>Going into the Lungs.</i>		<i>Coming out of the Lungs.</i>	
Oxygen.....	20.00	Oxygen.....	16.00
Nitrogen.....	79.96	Nitrogen.....	79.96
Carbonic acid gas.....	0.04	Carbonic acid gas.....	4.04
	100.00		100.00

It will be seen that 4 per cent. of the oxygen is converted into carbonic acid gas, so that air once breathed contains 100 times more of the gas than pure air, and since the carbonic acid gas is poisonous if inhaled in quantities of 4 per cent., it becomes apparent how essential it is to have good ventilation in our houses. Air as exhaled is wholly unfit and dangerous to breathe again. The air in houses does not contain carbonic acid gas from this source only—lamps

and gas when burning produce a considerable quantity, a gas flame, for instance, as much as five or six persons.

It would seem that the quantity of this gas in the atmosphere (0.04) would gradually and constantly increase, but Nature, in its wise economy, provides that this gas, which is noxious to animal life, should serve an important rôle in the sustenance of the vegetable kingdom, in that it furnishes the greater bulk of the food of all plants. Plants draw their nourishment from two sources—from the soil, by absorption of moisture and mineral matter, and from the atmosphere, by *breathing the carbonic acid gas, retaining or fixing* the carbon in their structure and *exhaling the oxygen*. The refuse of one kingdom is the essential life-principle of the other—animals breathe oxygen and produce carbonic acid gas; the plant breathes this carbonic acid gas, resolves it into carbon, which it retains, and oxygen, which it gives out again—a beautiful natural reciprocity.

The other constituents in the air are of minor importance, and their proportion is never constant. Small proportions of nitric acid and ammonia are always present, while sulphurous acid and sulphuretted hydrogen are only found in the air of crowded cities or near manufactories, etc.

Ammonia may sometimes be detected in rain water, which washes it out of the air, while nitric acid is usually found after thunderstorms.

The latter is thought to be formed by electricity, every flash of lightning causing the oxygen and nitrogen in the air to combine with some of the watery vapor to form the acid. Such constituents of the air as the odorous emanations from flowers, the various bacteria, saline particles

from the spray of the ocean, dust, etc., are only found at times, and are restricted to localities. The moisture in the air varies much, sometimes being present in as small a quantity as $\frac{1}{2}$ per cent. or even less, while at other times as much as 2 per cent. is present, when the air is wholly saturated. We feel so uncomfortable on warm, close days because the air is so saturated with aqueous vapor that the exhalations from our bodies (perspiration) cannot be taken up by it and consequently cling to our bodies.

Carbon.

Occurrence in Nature.—Carbon occurs in nature as coal, graphite, plumbago, diamond, etc. The diamond and graphite are crystallized forms, while all, or most all, other forms are amorphous. In combination it is the essential constituent of all organic matter; indeed, the chemical study of organic bodies is now termed “the chemistry of the carbon compounds,” while all organic bodies are included in the general term *carbon compounds*. Carbon is also a constituent of carbonic acid gas; all carbonates contain it, as well as all cyanides. We are familiar with carbon also as lampblack, bone-black, charcoal, gas carbon, coke, soot, etc. We thus see that this element occurs in several forms, the three principal ones being the diamond, graphite and coal.

This property of existing in two or more conditions or forms, which are distinct in their *physical* or *chemical* relations, is called *allotropism*, and the forms *allotropic* forms. There is no explanation for the existence of these conditions. Why the diamond occurs crystallized in octahedral (eight-sided) and other related forms in a state of extreme hardness, while blacklead occurs in hexagonal

(six-sided) forms, with very little hardness, and lampblack with no hardness at all and devoid of crystalline shape, we are at a loss to explain. Some chemists claim that these allotropic forms represent the *passive* and *active* states of carbon. (Ozone is an *active* state of oxygen and distinct from the ordinary oxygen, which is the element in its *passive* state.)

PROPERTIES.—Carbon in any form is insoluble, not fusible nor volatile, but combustible, forming carbon monoxide or the dioxide usually. Chemically it combines with many of the non-metals, but principally with hydrogen, with which it forms that important class of bodies known as *hydrocarbons*; with oxygen to form two oxides, CO and CO₂, and with nitrogen in the organic kingdom.

With hydrogen alone carbon forms *hydrocarbons*, but with hydrogen and oxygen, and the latter two in the proportion to form water, it forms *carbohydrates*. CH₄, marsh gas, is a hydrocarbon. C₁₂H₂₂O₁₁, cane sugar, is a carbohydrate. The product of the combustion of carbon is always carbonic acid gas, CO₂, at times accompanied with carbon monoxide, CO.

The *diamond*, the most precious of the gems, is the purest form of carbon. Its hardness and refractory power are characteristic and give it its value; it is one of the hardest substances known, and when properly cut refracts and resolves white light into the primary colors like no other body. When heated intensely it burns, producing CO₂, but resists the action of any known chemical.

PLUMBAGO, or graphite, is an almost pure carbon, though most specimens contain a small proportion of iron. We are all familiar with this form of carbon in lead pencils, the graphite being erroneously called *lead* at times.

LAMPBLACK is the soot produced when oil or some kinds of resin burn with an insufficient quantity of oxygen. It contains unchanged oil frequently.

CHARCOAL is produced when wood is incompletely burned, usually by a process known as destructive distillation, as in the preparation of acetic acid. It may be prepared also by the smothered combustion of piles of wood partially covered with earth. This kind of charcoal is called *wood charcoal*, to distinguish it from *animal charcoal*, which is the residue obtained on heating bones to a red-heat without access of air. This residue contains, besides carbon, phosphate of calcium, which when removed with hydrochloric acid leaves the *purified animal charcoal*. This is a very valuable substance, because of its property of removing coloring matter from solutions of organic matter. It plays an important part in the refining of sugar. The crude juice from the sugar-cane yields upon evaporation a very dark-brown sugar, whose color is due to natural coloring matter. The white granular or block sugar is the same devoid of the coloring matter, which is removed by filtering the solution of brown sugar or the crude cane-juice through animal charcoal contained in large percolators.

The manufacturing chemist employs animal charcoal for similar purposes, and for removing certain other bodies from solution. Nearly all varieties of charcoal possess this property, but none so much as the animal charcoal.

Charcoal prepared from dense wood has the property of condensing gases and vapors; it will absorb 90 volumes of ammonia gases, while of hydrogen it takes up about twice its own volume. In thus absorbing the gas the charcoal probably causes the gases to liquefy.

Compounds of Carbon with Hydrogen.—Carbon and hy-

drogen form compounds which are the starting points in *organic* chemistry. From these *hydrocarbons* a vast number of compounds are derived, either by adding, taking away or substituting other atoms or groups of atoms; indeed, the complete study of the hydrocarbons, their compounds and substitution-derivatives, constitutes the entire science of organic chemistry. These bodies are obtained in nature usually by the decomposition of more complex organic substances, but they may be made by direct combination, in many cases, in the laboratory. Acetylene, C_2H_2 , for instance, may be obtained by directly combining the elements by passing a stream of hydrogen through a tube in which a current of electricity is passing between two carbon poles.

Methane, or marsh gas, CH_4 , is another hydrocarbon of simple construction, and is produced when organic matter decays under water, as it does in marshes, hence the name. It, too, is the beginning of a long series of carbon compounds, and with the acetylene and ethylene or olefiant gas, is an important factor.

Illuminating Gas.—The manufacture of coal or illuminating gas is an industry of importance. The process is a simple one, yet in practice requires considerable skill. Coal is subjected to destructive distillation, which produces a variety of gases and other products, among which is a considerable quantity of ammonia, from the nitrogen which is always present in coal. The distillation of the coal goes on in large cast-iron retorts maintained at a red heat in a furnace. The gases and all the volatile products are conducted upward into and through a long horizontal pipe, containing water and called a hydraulic main. From there the gas passes into an apparatus which serves the

purpose of a refrigerator, consisting of a series of vessels filled with water and so arranged that the gas must pass through a considerable quantity of water. The procedure not only cools the gas but causes the condensation of tar and ammonia. Much tar is deposited in the hydraulic main, from where it runs into a reservoir known as the *tar-well*, but the gas being still very hot not all the tar condenses, and some passes over, to be deposited in the refrigerator with the ammonia. The process of purification is continued by causing the stream of gas to pass through a chamber called the *scrubber*, which is filled with loose material, or anything that gives space. Here much suspended matter is removed. Next the gas is *washed* by being made to pass through an apparatus where it comes in contact with a number of sprays of water. The gas still contains sulphuretted hydrogen, carbonic acid gas and disulphide of carbon, which are removed in the *purifiers*. The latter are very large tanks filled with lime. The impurities mix chemically with the lime and are retained. The gas is finally stored in the large gas tanks, with the appearance of which we are all familiar, and from which it is supplied to the consumer.

Illuminating gas varies much in its composition, judging from its illuminating powers. If properly purified it contains methane, CH_4 ; ethylene, C_2H_4 ; acetylene, C_2H_2 , hydrogen, carbon monoxide, CO ; free nitrogen, a little disulphide of carbon and a small quantity of volatile liquid hydrocarbons.

Compounds of Carbon with Hydrogen and Oxygen.—The *carbohydrates*, as already observed, are a class of compounds containing *carbon*, united with *hydrogen* and *oxygen* in the proportion to form water; there are usually six

carbon atoms or a multiple of six present. There are three groups of these compounds: the *glucoses*, represented by grape sugar, fruit sugar, and all having the composition $C_6H_{12}O_6$; the *saccharoses*, of which cane sugar, milk sugar and maltose are examples, having the composition $C_{12}H_{22}O_{11}$, and finally the *amyloses* (starches), $C_6H_{10}O_5$, with starch, dextrin and cellulose as representatives. It will be seen that the hydrogen and oxygen in each case are in the proportion to form water—in cane sugar, for instance the $H_{12}O_6$ are equal to $6 H_2O$ —i.e., six molecules of water. Cellulose, of which ordinary absorbent cotton is a good example, has a formula represented by a multiple of the six carbon atoms usually present in the carbohydrates; its formula is that of *amylose* or starch multiplied by three— $C_{18}H_{30}O_{15}$. The carbohydrates, like the hydrocarbons, are an exceedingly interesting and useful class of compounds, but cannot be considered here.

Alkaloids and Glucosides.—Carbon united with *hydrogen* and *nitrogen* alone, or with these and oxygen, forms the two series of compounds known as *alkaloids* and *glucosides*. These, too, are relegated to organic chemistry and cannot be dwelt upon here.

Compounds with Oxygen.—There are two compounds of oxygen and carbon: CO, carbon monoxide, and CO_2 , carbon dioxide.

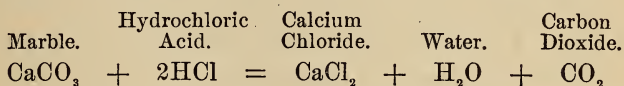
CARBON MONOXIDE is a combustible gas, producing carbon dioxide when it burns: $2CO + O_2 = 2CO_2$. Its flame is a beautiful pale blue, which may be viewed in any coal stove when the coal has just begun to burn. The gas is without color, almost without odor, and is very poisonous. It is produced abundantly, together with carbon dioxide, when charcoal burns. It rapidly produces in-

sensibility and death when inhaled. When oxalic acid is heated with five or six times as much sulphuric acid, carbon monoxide and dioxide are formed, from which mixture the dioxide may be removed by passing through a strong solution of caustic potassa. The dioxide unites with the potassa to form potassium carbonate, while the monoxide passes unchanged. Another method for the production of the monoxide is the heating of potassium ferrocyanide with ten times its weight of sulphuric acid. An abundance of the gas in a pure state is thus generated. Extreme care must be observed not to inhale this gas.

CARBON DIOXIDE, sometimes erroneously termed *carbonic acid*, is always produced when anything containing carbon burns with a sufficient supply of oxygen. When the supply of oxygen is insufficient the monoxide is often formed. In these two gases we have a good illustration of the law of multiple proportions referred to in the paragraphs on chemical philosophy. If carbon, whose atomic weight is 12, combines chemically with oxygen, and there is only a limited supply of the latter, the combination takes place between 12 of the carbon and 16 of the oxygen, forming 28 of CO, but if the oxygen is plentiful the carbon unites not with once the atomic weight of oxygen, 16, but with a multiple of 16—namely, 32, producing 44 of CO₂; that is, while the proportion of carbon is constant, that of oxygen increases, in the case of CO₂, in a multiple proportion. *Combustion* is ordinarily *oxidation*—that is, the union, chemically, of a body with oxygen, with the production of heat and flame. That body may be hydrogen, sulphur or phosphorus, etc., but usually it is carbon in some form or other. The burning of coal, wood, gas or oil is simply the union of the carbon in those bodies

with the oxygen in the air, producing carbon dioxide. The blood is continually purified in our lungs by the carbon in the refuse matter collected by the blood uniting with the oxygen brought into the lungs as air.

For experimental purposes the gas is easily prepared by decomposing any carbonate (that of sodium is cheap) with an acid. On a larger scale, as in the making of the carbonated waters, marble, the carbonate of calcium, is decomposed with acid, hydrochloric usually, or sulphuric:



Any carbonate when treated with one of the strong acids yields free CO_2 . This fact is taken advantage of in testing for carbonates. Whenever an effervescence (evolution of gas) takes place when a body not a metal is treated with an acid, it is safe to assume that the body was a carbonate.

Carbon dioxide is a colorless gas, with an agreeable pungent taste and odor. It is poisonous to animal life, as already observed when discoursing on the atmosphere. It is incombustible and does not support combustion—a lighted match or taper plunged into the gas is immediately extinguished. In this respect and in others it is not unlike nitrogen, but may be distinguished from it by its rapid absorption by solution of caustic potassa or lime-water. The latter with the gas becomes turbid, owing to the formation of insoluble carbonate of calcium. The gas can thus be detected in the breath by blowing through a tube into lime-water.

Water dissolves about its own bulk of CO_2 , but many times its volume if under pressure. Soda-water is common water charged with carbon dioxide under pressure. The

soda-water can be drawn through pipes upward for a distance, because the gas, owing to its expansive force and pressure, tries to get away from its confinement and carries the water with it. Effervescing wines and *sparkling* mineral waters owe their effervescing qualities to carbon dioxide. In the former it is produced during the process of fermentation, when the grape juice changes its sugar into alcohol; in the latter it is present through the absorption of the gas produced in nature, usually in the earth near springs. It is always a product of *fermentation* and *putrefaction*. These latter terms are synonymous, in a sense—they both mean the decomposition of organic matter. When the products of the decomposition are useful the process is termed fermentation, as in the formation of alcohol by the decomposition (natural decomposition) of grain; but when the products are useless and offensive the process is called putrefaction. In many of the mineral waters carbon dioxide plays an important part, as has been mentioned in the chapter on water, to which the student is again referred.

Carbonic acid gas, as the gas is often correctly termed, can be liquefied; it requires for the purpose a pressure of 577 pounds to the square inch. When liquefied it is colorless, lighter than water and, when unconfined, four times more expansive than air. Because of this and its power to extinguish fire it may be found in many places kept in strong glass flasks hermetically sealed, to be used in case of fire. The flask breaks when thrown into a burning mass; the liquid at once volatilizes, extinguishing all flame it comes in contact with. Care must be taken that no persons or animals are near.

THE CARBONATES form a very large and important

series of salts, many of which occur abundantly in nature. Carbonate of calcium is found in nature in many forms, most largely as marble, limestone and chalk. The shells of all crustaceous animals, such as oysters, clams, lobsters, etc., are wholly composed of calcium carbonate. The carbonate of magnesium also occurs very abundantly. The carbonates may be represented as containing the oxide of a metal united with CO_2 . Chalk, for instance, CaCO_3 , is CaO , oxide of calcium, and CO_2 ; carbonate of magnesium, MgCO_3 , is the oxide, MgO , and CO_2 ; white lead, PbCO_3 , is the oxide of lead, PbO , and CO_2 .

CARBONIC ACID, H_2CO_3 , may be looked upon as water, H_2O , and CO_2 . The compound is not stable and always resolves itself into water and CO_2 . Attention should here be called to an error which is too often made. CO_2 is *carbon dioxide*, or carbonic acid gas, and not *carbonic acid*, as it is so frequently called. There can be no acid without hydrogen, and *carbonic acid* is CO_2 with water, H_2O , which furnishes the hydrogen.

Carbon with the Halogens.—There are a number of compounds of carbon and chlorine known: C_2Cl_2 , C_2Cl_4 , C_2Cl_6 , etc., and a similar group with iodine and bromine.

Carbon with Nitrogen forms *cyanogen*, CN , which is a colorless gas of a pungent, peculiar odor resembling somewhat that of hydrocyanic acid. Cyanogen unites with hydrogen to form the most poisonous compound there is, *hydrocyanic acid*, sometimes called *prussic acid*. With metals it forms compounds called *cyanides*. Cyanogen is a good illustration of what is meant by a *compound radical*; it is a group of elements, carbon and nitrogen, which combines with elementary bodies, and is capable of passing from one state of combination to another just as though

it were a single element. Ammonium, NH_4 , is analogous to cyanogen in this respect, except that it acts as a base or like the metals, while cyanogen is related to the non-metals in the same sense.

Carbon with Sulphur forms, like carbon and oxygen, two compounds, the monosulphide, CS , and the disulphide, CS_2 . The disulphide is of some importance in that it is a good solvent for some bodies which are ordinarily insoluble. It has a very offensive, repulsive odor as found in the market, but when pure is almost odorless. Sulphur is easily dissolved by it, and is again deposited in crystals upon evaporation of the disulphide. Its use is mostly in the arts, though to some extent it is utilized in extracting oil from seeds. It dissolves or extracts the oil, leaving it again upon evaporation or distillation.

The Halogens.

THE four elements known as the *halogens* (salt producers, so called because they exist in sea-water) are chlorine, bromine, iodine and fluorine. We will consider them in the order named and begin with

Chlorine.

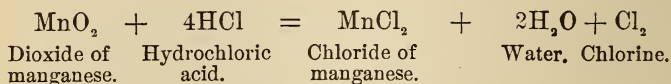
CHLORINE is one of the natural group of elements called halogens, which not only have a common source, but resemble each other in a marked degree in all their chemical relations. The similarity in chemical behavior is especially noticeable between chlorine, bromine, and iodine; fluorine, because chemists have not been able to isolate it in any quantities large enough to experiment with, has not been so thoroughly studied, but its behavior in combination is very like that of the other halogens in many respects.

Occurrence in Nature.—The halogens occur chiefly in the sea, usually in combination with sodium, potassium, magnesium and calcium. Chlorine is the most abundant, and is found mainly as the chloride of sodium—common salt—in the ocean, as brines, and in mines as solid crystalline masses. Some springs contain it; the St. Catherine's water contains it in combination with magnesium. As the chloride of sodium it is universally distributed in minute quantities; the tissues of plants and animals contain it as an essential constituent.

History.—The Swedish chemist, Scheele, who also discovered oxygen, discovered chlorine. While experimenting with hydrochloric acid he brought it while hot in contact with *braunstein*, brownstone, dioxide of manganese, when he noted an irritating, suffocating gas coming off, which he and others subsequently proved to be an element; this also was in 1774. It is a peculiar coincidence that oxygen should have been discovered at about the same time by two men, Priestley and Scheele, distant from one another and without knowledge of each other, but it is more so that one body—dioxide of manganese—should have served in the discovery of two elements, for Scheele not only discovered oxygen by heating the dioxide, when oxygen was evolved, but chlorine as well by treating the dioxide with hydrochloric acid. In the first instance the element, oxygen, was a constituent of that from which it came, MnO_2 , but in the case of chlorine the dioxide was only an agent for the liberation of chlorine from hydrochloric acid, and did not furnish the chlorine itself.

Preparation.—The accidental method which furnished Scheele's chlorine is to-day yet employed as the simplest, though there are many others. Manganese dioxide, MnO_2 ,

is introduced into a flask provided with a safety-tube for pouring in the acid and a bent tube for conveying the gas to a receiver. A small portion of hydrochloric acid is poured through the safety-funnel tube into the flask and thoroughly shaken, so that all of the manganese becomes wetted; then more acid is added and gentle heat applied, when the evolution of the gas ensues at once, and may be collected as is any other gas. The equation illustrative of this reaction is:



Instead of using hydrochloric acid, sulphuric acid and common salt are sometimes employed—i.e., the dioxide is mixed with common salt, to which mixture the sulphuric acid is added. The common salt and sulphuric acid react upon each other, producing hydrochloric acid: $2\text{NaCl} + \text{H}_2\text{SO}_4 = \text{Na}_2\text{SO}_4 + 2\text{HCl}$. The HCl thus produced in the operation is then acted upon by the manganese dioxide with the result above illustrated.

Chlorine gas, thus produced and passed into water, constitutes the *aqua chlori* of the pharmacopœia, which contains 0.4 per cent. of gas. It does not keep well, producing hydrochloric acid through the decomposition of the water by the union of the chlorine with its hydrogen. Keeping in dark places prevents this change to an extent, but the water is never reliable if kept for any length of time.

Properties.—In chemical energy chlorine surpasses many of the elements, even oxygen in some cases. It has a greenish color which is characteristic, a peculiar suffocating odor, and when inhaled produces distressing irritation of

the mucous membranes of the throat and lungs. Some claim that if respired in very minute quantities it is beneficial in pulmonary troubles. A pressure of four atmospheres liquefies the gas to a transparent yellow liquid which it is difficult to freeze into solid condition. That chlorine maintains combustion may be shown by dropping powdered antimony, arsenic or phosphorus into it; the chlorides of these will be formed almost immediately, with the production of heat and flame. Burning is ordinarily *oxidation*; in this case it is *chlorination*. An interesting experiment may be made by saturating a piece of paper with oil of turpentine and plunging it carefully and dexterously into a vessel containing chlorine. A dense black smoke is emitted which usually turns into a flame from the rapid decomposition of the turpentine.

The atomic weight of chlorine is 35.4, its density the same, and its specific gravity compared to air 2.45. It has but little attraction for oxygen, but all the more for hydrogen and the metals, with which it forms chlorides. Wood or wax will burn in chlorine with a dull red flame, burning the hydrogen only, producing hydrochloric acid. The bleaching power of chlorine is its most characteristic property; organic coloring principles which resist the bleaching action of other bodies are almost instantly destroyed by it. Indigo, which will not change its color with strong sulphuric acid, has its color discharged by chlorine. Whenever chlorine bleaches it does so not by hiding, so to say, the color, as sulphur dioxide does, but by completely destroying it, so that it cannot be restored. Water must be present for chlorine to act as a bleaching agent, as the gas in a perfectly dry condition is devoid of any bleaching qualities; it will not even affect

litmus when dry. In pharmacy and in the arts and industries chlorine is used largely in the form of *bleaching-powder*, which contains 35 per cent. of available gas. It is employed principally for bleaching cotton and linen goods. Chlorine is one of the most thorough and potent disinfectants, and for this purpose is usually employed to form the bleaching-powder, or chloride of lime, as it is erroneously termed—*chlorinated* lime is the proper term. When chlorinated lime is mixed with water in a flat vessel it decomposes slowly through the action of the carbonic acid gas in the air, setting the chlorine free. Mixing a little hydrochloric acid with it serves to disengage the gas more rapidly. It is not prudent to use the gas for disinfecting purposes, its use by inexperienced persons being fraught with danger.

Compounds Containing Chlorine.—With *oxygen* chlorine cannot be united directly, all the known combinations being produced indirectly. All the oxides and corresponding acids decompose very easily, sometimes with *violent explosion*, for which reason many of their compounds are employed in making explosives. The oxides and acids obtained from them by their union with water are the following:

Cl_2O , *hypochlorous oxide*, $+ \text{H}_2\text{O} = 2\text{HClO}$, *hypochlorous acid*.

Cl_2O_3 , *chlorous oxide*, $+ \text{H}_2\text{O} = 2\text{HClO}_2$, *chlorous acid*.

Cl_2O_4 , *tetroxide*. Acid not known.

Cl_2O_5 , *chloric oxide*, $+ \text{H}_2\text{O} = 2\text{HClO}_3$, *chloric acid*.

Cl_2O_7 , *perchloric oxide*, $+ \text{H}_2\text{O} = 2\text{HClO}_4$, *perchloric acid*.

In viewing the composition of these oxides it will be seen that the quantity of chlorine is constant in each, the quantity of oxygen increasing, which illustrates again the

law of multiple proportions. It not only illustrates this law, it teaches also

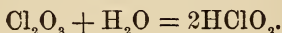
A Lesson in Nomenclature,

i.e., the naming of compounds. Here we have a series of combinations, all containing the two elements, but each containing one element in different proportion from that in another. Now, it is this difference, in the quantity of oxygen in this case, that enables chemists to apply a definite specific name to each combination. The terms mono, di, tri, tetra and pentoxide would be names for combinations containing, respectively, one, two, three, four and five atoms of oxygen. Besides these numerical terms the chemist employs others which are *relative* ones, the relation starting from some one combination. These *relative* terms are the true *chemical* names. To the highest form of oxidation is usually given a name, to which the names of the others bear a relation. In this case the Cl_2O_6 was once thought to be the highest form of oxidation of chlorine, and was called chloric oxide, the *ic* always meaning *highest*. The Cl_2O_5 was the next highest, or the one next below the highest, and it was termed chlorous oxide, the *ous* meaning *lower*. There being also a Cl_2O , that is, with the Cl_2O_5 , two chlorous oxides, another prefix had to be employed to show that the Cl_2O was a combination of the *ous* kind, though lower than one already called chlorous. The prefix *hypo* was called into requisition, which means *beneath*, so that the *hypochlorous* oxide Cl_2O is a combination *lower* than the chlorous oxide. For the Cl_2O_7 , a name had to be chosen which would indicate that it was a union of chlorine and oxygen with even more oxygen in

it than the one termed *chloric oxide*. The prefix *per* means *above*, so that the word *peroxide* means *above the highest oxide*. (Intrinsically the Cl_2O_7 would be the *chloric oxide*, etc.) What has been said of the names of the combinations of oxygen and chlorine is not confined to these, but the principles are *generally applicable* in chemical nomenclature.

With *hydrogen*, and with *hydrogen* and *oxygen*, chlorine forms *acids*, which may be directly derived from the above oxides by adding water, except in the case of hydrochloric acid, as can be seen above. Most oxides of the non-metals form acids with water, the hydrogen in the water forming the base of the acid, and without which no acid can exist.

It will be seen that with *water* the various non-metallic oxides form *acids* which have the same names as the oxides. Knowing, therefore, the name of an oxide, we can tell the name of its corresponding acid, and *vice versa*. The *corresponding acid* of an oxide is what is obtained when water is added to that acid; for instance, the acid corresponding to the chlorous oxide, Cl_2O_3 , is chlorous *acid*:



The names of the acids furnish us names for the salts which are formed from the acids. Salts are generally derived by replacing the hydrogen in acids by a metal. Thus, if we replace the H in hydrochloric acid, HCl , with sodium, Na, we get NaCl , which is a salt; so by putting an atom of magnesium, Mg, in place of the H in sulphuric acid, H_2SO_4 , we get MgSO_4 , also a salt. Let us now make *salts* from all of the chlorine acids. The only acid not containing oxygen in this series is hydro-

chloric acid, which can be made by direct union of H and Cl.

If in the following acids the hydrogen is replaced by sodium,		The following-named salts will be produced :	
Hydrochloric acid, HCl...		{ Chloride of sodium, or sodic chloride.....	{ NaCl
Hypochlorous " HClO.		{ Hypochlorite of sodium, or sodic hypochlorite.....	{ NaClO
Chlorous " HClO ₂		{ Chlorite of sodium, or sodic chlorite.....	{ NaClO ₂
Chloric " HClO ₃		{ Chlorate of sodium, or sodic chlorate.....	{ NaClO ₃
Perchloric " HClO ₄		{ Perchlorate of sodium, or sodic perchlorate.....	{ NaClO ₄

It will be seen from this that acids having the prefix *hydro* and the suffix *ic* make salts ending in *ide*: *hydrochloric* acid yields a *chloride* of any metal. The acid whose name ends in *ous* and has *hypo* for prefix yields salts called *hypochlorites*; the *ous* acids yield *ite* salts: chlorous acid gives *chlorite* of a metal, sodium, in this illustration; the *ic* acids yield salts whose names end in *ate*: *chloric* acid yields *chlorates*; acids having names with *per* for prefix and *ic* for suffix yield salts whose names begin with *per* and end in *ate*: *perchloric* acid yields *perchlorates*, it being understood that in each case the hydrogen of the acid is replaced by a metal.

Experiments.—Before beginning the discourse upon the uses of chlorine in pharmacy, directly and indirectly, its importance warrants the giving of a little time to experimenting, especially with a view of illustrating more fully its properties, chemical and physical. Besides those given before, the following experiments may be made by the student. He is advised to experiment whenever practicable, and to record his observations for future reference, and always to be *careful* in his experiments.

1. ILLUSTRATING DIFFUSION OF GASES.—If chlorine gas be made, as before directed, by displacing water in a flask and the flask exposed to the atmosphere, it will take a long time for the green color of the gas to disappear; the disappearance of the color naturally indicates the disappearance of the gas. If another flask containing chlorine gas be inverted, the flasks being in both instances without stoppers, it will require only a very short time for the chlorine to mix with the atmosphere and the flask to become filled with air.

This phenomenon is due to the fact that chlorine is 2.5 times heavier than air. If it were not for a natural law, the *law of diffusion of gases*, the chlorine in the first flask would remain there indefinitely, being much heavier than air, for it must be remembered that gases, like liquids, when carefully mixed, arrange themselves in layers according to their densities. The layers of liquids (in case of miscible liquids) when once arranged remain so permanently; for instance, a test-tube containing sulphuric acid may have a layer of water poured in without disturbing the sulphuric acid—i.e., the acid and water form two layers which remain permanent if not disturbed by some external force. The same is true of gases, so far as their arrangement according to their density or heaviness is concerned, *at first*, but the layers are not permanent. Notwithstanding the different densities of gases, they will diffuse or intermingle with each other in admixture until the mixture will have attained a mean density. Where there is no limit by a containing vessel, as in the atmosphere, the diffusion practically goes on indefinitely. (See physical properties of hydrogen on pages 29 and 30.)

2. ILLUSTRATING BLEACHING PROPERTIES.—Make chlorine gas as before, and pass it through a tube or bottle containing chloride of calcium, and conduct it to the bottom of the bottle which is to contain it. The bottle will become filled with chlorine, which *displaces* the air. The object of passing the gas through the calcium salt is to remove any watery vapors with which it may be mixed, to obtain a thoroughly *dry* gas. The chloride of calcium has a great affinity for water, and absorb it from its mixture with a gas. A piece of dry litmus-paper or a dry red rose plunged into this dry chlorine gas will not have its color affected. Moisten the litmus-paper or the rose, and again introduce into the chlorine. The color begins to discharge, and very soon the paper or rose will be perfectly white. This illustrates that chlorine will not bleach unless water is present.

A little chlorine-water poured into ink or red wine or indigo solution causes each of the liquids to lose its color. *All colors derived either from the animal or vegetable kingdom are bleached or destroyed by chlorine.* This property makes chlorine a most important agent for bleaching cotton, linen, paper and many other materials. It renders these bodies perfectly white in a few hours, while the old method of bleaching by laying the goods on the grass in sunshine required weeks, and sometimes months.

3. ILLUSTRATING DISINFECTING AND DEODORIZING PROPERTIES.—If chlorine-water is applied to decaying and badly smelling substances (stagnant water, rotten eggs, etc.) the bad odor vanishes very quickly.

It not only decomposes colors, but also the products formed during decay, and which produce nauseous odors. Morbific matter, disease germs, miasms, which, by being

diffused in the air or attached to beds, clothes or walls may communicate disease, are acted upon in a similar manner. Chlorine is therefore a deodorizer and powerful disinfectant, purifying infected atmospheres and all kinds of morbid matter, and arresting or checking the decay of organic substances.

4. ILLUSTRATING THE AFFINITY OF CHLORINE FOR HYDROGEN.—Fill a small flask with chlorine-water and invert into a vessel containing water, and put away into the dark; there will be no change and the chlorine water will keep for a tolerable length of time; but if the flask be exposed to sunlight decomposition will soon ensue, with the formation of a gas which collects in the upper part of the flask. After a few days the water will have lost all of its chlorine odor and will have acquired a sour taste, and instead of bleaching litmus paper will turn it red, showing acid reaction. The gas collected in the upper part of the flask will be found to inflame a glowing taper and exhibit other properties which establish its identity as oxygen; the acid reaction is due to the presence of hydrochloric acid. We began with chlorine and the elements of water, hydrogen and oxygen. These, later on, are yet present, but not in the original condition or combination—the chlorine being now not free or uncombined, as it at first was, but in combination with the hydrogen as hydrochloric acid. In other words, chlorine has much chemical affinity, so much, in this case, that it decomposes water, under the influence of sunlight, and unites with the hydrogen, setting free the oxygen, thus: $2\text{Cl}_2 + 2\text{H}_2\text{O} = 4\text{HCl} + \text{O}_2$. When the water is decomposed the chlorine has the choice of uniting with oxygen or with hydrogen, but it selects the latter, affording an example of what is termed *simple elec-*

tive affinity. Now, this very chemical affinity of chlorine for hydrogen explains its bleaching and disinfecting power. When chlorine bleaches it simply causes the decomposition of the coloring matter in order to unite with the hydrogen which is present in all animal and vegetable bodies. By this abstraction of the hydrogen a rearrangement of molecules to form other combinations is necessary, causing the coloring matter to become colorless.

5. ILLUSTRATING THE AFFINITY OF CHLORINE FOR BASES.—Make a solution of iodide of potassium or of the iodide of any base, and add freshly-made chlorine-water; the iodine will be set free and the chlorine take its place, making a chloride of potassium.

Treat a solution of any bromide with chlorine-water or any form of free chlorine—the bromine becomes liberated and may be recognized by its odor and color. This shows the chlorine to have more affinity for the base than the bromine or iodine has, which it replaces in all their combinations.

6. SHOWING THE SOLVENT POWER, CHEMICALLY, OF CHLORINE.—Into chlorine-water put a leaf of *pure* gold; it will disappear in a short time, uniting with the chlorine to form chloride of gold. It is the only substance that will dissolve gold (nitro-hydrochloric acid dissolves gold, but it does so because it probably contains free chlorine). Chlorine combines thus with many of the metals, forming *chlorides* of the respective metals, as has already been referred to in case of antimony and arsenic. Copper in the form of wire or sheet burns in chlorine gas, forming a solution of copper chloride. If into this latter solution a piece of bright iron or steel is placed it very soon becomes covered with a red coating of copper. In chemical reac-

tions the right of the strongest prevails, and the chlorine, possessing stronger affinity for iron than for copper, seizes the iron, forming a chloride of iron, and liberates copper, which is deposited in the metallic condition. Polished steel is, therefore, a reagent for copper and may be employed to ascertain the presence of copper very simply and quickly. Copper is sometimes detected in this way if present in pickled cucumbers or preserved fruit, to which it is sometimes added to give a fresh-green color.

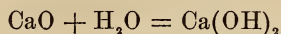
7. SHOWING THE TEST OF IDENTITY OF CHLORINE IN COMBINATION AS CHLORIDE.—To a solution of any chloride add solution of silver nitrate—a heavy white precipitate will be formed which is insoluble in nitric acid, but quickly soluble in ammonia-water.

CHLORATES when heated evolve oxygen, the residue being chlorides.

Chlorine in Pharmacy.—*Aqua chlori*, U. S. P. has been already incidentally considered.

Calx Chlorata.—This is a compound resulting from the action of chlorine upon calcium hydroxide (slaked lime), and containing at least 35 per cent. of available chlorine. It is erroneously called chloride of lime—properly it should be termed *chlorinated lime*. Calcium hydroxide, $\text{Ca}(\text{OH})_2$, placed in trays in suitable chambers, is exposed to the action of chlorine, which is absorbed by the lime, forming a compound having the formula $\text{CaCl}_2\text{Ca}(\text{ClO})_2$. Its formula is sometimes written CaOCl_2 (the former divided by two), which gave rise to the name *chloride of lime*, the CaO being the formula for *lime* and the Cl_2 denoting *chloride*. In correct chemical nomenclature we cannot recognize the latter name. It is proper, though, to speak of the compound as *chlorinated lime*, the hydroxide

being made from lime, CaO , by adding water or slaking it:



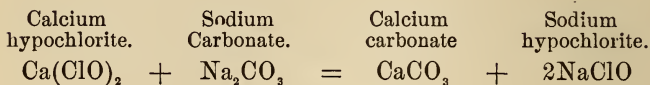
Chlorinated lime is valuable chiefly for the chlorine it contains, and which is easily available, being very loosely combined. It is used largely as a disinfectant through its power of arresting decay and putrefaction. Sometimes, though rarely, it is used internally. Its chief use in pharmacy is to make the solution of chlorinated soda of the Pharmacopœia, or Labarraque's Solution.

The manufacture and introduction into the market of chlorinated lime, or *bleaching-powder*, as it is frequently termed, was brought about by manufacturers being prohibited from allowing the chlorine which was a by-product in many processes of manufacture, to escape into the air. Air containing chlorine is not fit for breathing, nor is water containing chlorine fit for fishes to live in, as the manufacturers soon learned when they tried to dispose of their waste chlorine by conducting it into the rivers. At a loss what to do with it, they began experimenting, and soon a chemist suggested passing it into lime. His suggestion was adopted readily, but it was not thought that this apparently useless product would soon be in such demand that more than waste chlorine would be necessary to make enough to supply the need for it. Exposed to the air it emits chlorine spontaneously; if it is desired to have the chlorine evolved rapidly, a little hydrochloric acid added will hasten the disengagement of the gas.

Bleaching-powder should be preserved in well-closed vessels and kept in dry places.

Liquor Sodæ Chloratæ.—The solution of *chlorinated* soda is another convenient form for the use of chlorine.

It is made by adding solution of common washing-soda, sodium carbonate, to solution of bleaching-powder. The calcium is replaced by sodium:



The calcium carbonate is insoluble and settles to the bottom, when the clear soda solution can be poured off.

If potassium carbonate is employed in place of the sodium salt a corresponding potassium hypochlorite will be formed, the solution of which is known as *Javelle's water*. The values of both these solutions are equal.

Hydrochloric and other acids containing chlorine will be referred to in the chapters on acids.

Iodine.

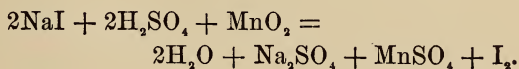
Occurrence in Nature.—Like other halogens, iodine occurs in nature chiefly combined with sodium and potassium; combinations with other bases are sometimes found. It chiefly occurs in the sea-water, but certain springs furnish it in large proportion.

Source.—Many marine plants have the power of absorbing or abstracting the combinations of iodine from sea-water and storing them in their tissues, and this source—marine plants—furnishes all of the iodine in commerce.

Preparation.—To obtain the iodine from the plants, the latter are burned, and the ashes, which are known as *kelp*, are lixiviated—that is, water is passed through them, as the menstruum is passed through the drug in the process

of percolation. The solution thus obtained is filtered and concentrated by evaporation, allowing it to cool at times, to permit the other less soluble salts to crystallize. The chloride and carbonate of sodium and the chloride of potassium—all salts found besides the iodides in the ashes of sea-plants—crystallize out in the order named, leaving the iodides, because most soluble, to remain until last in solution.

The dark-brown liquid which is left, called the *mother-liquor* (a mother-liquor is that liquid which is left, in any process of crystallization, after one or more crops of crystals have been separated), and which contains the iodine almost wholly as iodide of sodium, with some iodide of magnesium, is treated in a manner exactly analogous to that employed in making chlorine, viz.: it is treated with sulphuric acid and dioxide of manganese and gently heated in a retort, usually a leaden one, when the iodine *distils*, or better, *sublimes* over, and is condensed and collected in a receiver. Theoretically the operation is similar to that for obtaining chlorine in one case.



In this and many other chemical respects iodine closely resembles chlorine and the other halogens.

Of late years considerable iodine has been obtained in Chili from the mother-liquors obtained from the crystallization of nitrate of sodium, which is abundant there as an efflorescence on the soil.

Purification.—The iodine as thus obtained is in very impure condition, the main impurities being iodide of

cyanogen, or cyanide of iodine as it is frequently termed, chloride of iodine and water. To remove these, *resublimation* is resorted to. The crude iodine is first very gently heated in proper vessels to allow the more volatile chloride and cyanide to escape; increased heat then sublimes the iodine; water is removed by bringing in close proximity with freshly-burned lime; the lime abstracts the water: $\text{CaO} + \text{H}_2\text{O} = \text{Ca}(\text{OH})_2$.

Detection of Impurities.—When the iodine adheres to the interior of the bottle it does so because of the presence of water. Sometimes as much as 10 or 15 per cent of water is present, but its presence is injurious only in so much that it renders preparations made from it weaker than they should be. If iodine containing water is shaken with chloroform it does not give a clear and limpid solution, but a milky one, from which the water will separate if allowed to stand for some time.

Chloride of iodine, if present, gives to water shaken with the iodine a brownish color, a deep brown one if much is present. To detect the cyanide the pharmacopœia directs: If the iodine be removed from a dilute aqueous solution by agitation with disulphide of carbon, and, after the separation of the latter, some dilute solution of ferrous sulphate, with a trace of ferric chloride, be added, and finally solution of soda, and the whole supersaturated with hydrochloric acid, no blue precipitate should make its appearance.

Properties.—Iodine is in form of crystalline scales of a bluish-black color and metallic lustre. It is nearly five times as heavy as water, in which it is only very sparingly soluble; soluble in alcohol, ether, chloroform and disulphide of carbon. At ordinary temperatures it volatilizes

slowly; when heated it melts and rises as purple vapors, which are the heaviest aëriform vapors known. Applied to the skin iodine imparts a deep brown stain, which may be removed with ammonia-water.

Iodine precipitates most of the alkaloids, forming insoluble compounds with them. In chemical behavior iodine resembles chlorine, but its chemical behavior is weaker. Its symbol is I; its atomic weight, 126.85.

Compounds.—Iodine combines with many of the non-metallic elements and with most of the metals, forming iodides and iodates. With oxygen it forms principally the I_2O_5 and three acids—the iodous, iodic and periodic.

Of the iodides, those of potassium and sodium are most important in pharmacy, with the ammonium and iron following. The hydrogen iodide, hydriodic acid, and the iodides of mercury, too, are largely used. The iodine in these is loosely combined, wherefore the salts should be kept in well-stoppered bottles, away from sunlight and in a cool place.

Pharmaceutical Preparations.—The *Tincture of Iodine* is probably the most largely used preparation of iodine. It is made by dissolving 70 grammes of iodine in enough alcohol to make 100 cc., and should be made frequently and in small quantity, as the iodine reacts with the alcohol, especially when exposed to sunlight. It is used mainly externally, though at times it is given internally. Water decomposes the tincture, with a separation of the iodine. Next to the tincture in importance is the *Compound Solution of Iodine*, or Lugol's solution, which differs from the tincture in being a solution of iodine in a solution of iodide of potassium. Iodine alone is not soluble in water, but in the presence of iodide of potas-

sium it becomes readily soluble, and an aqueous solution is thus possible. Five parts of iodine and 10 parts of iodide potassium will dissolve in 85 parts of water.

This solution is intended mainly for internal administration, and its medicinal value depends not upon the iodide of potassium but upon the iodine, and the dose should be regulated according to it. There is no chemical reaction between the iodine and potassium iodide, the latter acting simply as a mechanical solvent.

The preparation of the *Ointment of Iodine* depends upon the same fact. Four parts of iodine are rubbed with one part of potassium iodide and two parts of water, and, finally, with enough benzoinated lard to make 100 parts. The object of the water and potassium iodide is to bring the iodine into a state in which it may be thoroughly and equally incorporated with the lard. Alcohol was formerly used to effect solution of the iodine, but the present method has been found to be a better one. A little glycerin obviates the formation of crystals in the ointment, which sometimes happens. If kept for a length of time this ointment undergoes change, hence it should be prepared when wanted for use. It is used whenever iodine is indicated and the ointment is the more convenient form of administration.

Iodized Starch is obtained by triturating starch and iodine with a little water and drying. It is used internally, but is not of much importance.

TESTS FOR IODINE AND IODIDES.—Free iodine gives a characteristic deep blue color with starch-paste. The latter is made by *boiling* a small quantity of starch with water.

Chlorine having more affinity for bases than iodine, the

latter can be liberated from iodides by treating their solution with chlorine-water.

Strong sulphuric acid dropped on any dry iodide also causes the iodine to be set free.

Any iodide gives, with a solution of silver nitrate, a yellowish-white precipitate, which is insoluble in nitric acid and but slightly soluble in ammonia-water. (Chlorides give a *white* precipitate, *readily* soluble in ammonia.) Mercuric chloride, or any soluble mercuric salt, gives, with an iodide, a red precipitate of mercuric iodide.

Any soluble lead salt produces a yellow precipitate with an iodide.

Bromine.

Occurrence in Nature.—Bromine, the third of the halogens, occurs in the sea and saline springs principally as bromide of calcium, sodium and magnesium.

Preparation.—It is chiefly obtained from *bittern*, the mother-liquor, obtained by evaporating *brine* from salt-wells. The mother-liquor containing the bromine, chiefly as magnesium bromide, is treated with chlorine, which, however, having greater chemical affinity for the bases, combines with magnesium, in this case, setting the bromine free. (It must be remembered that chlorine has the greatest affinity of the halogens.) The bromine is condensed and collected in cooled receivers under water. This equation expresses the reaction: $\text{MgBr}_2 + \text{Cl}_2 = \text{MgCl}_2 + \text{Br}_2$.

Properties.—Bromine is the only liquid non-metallic element. At ordinary temperatures it is a dark-brownish red, mobile liquid, giving off brown fumes of an exceedingly offensive and suffocating odor; it is very volatile

and must be kept under water, in which it is not very soluble. It is freely soluble in ether, carbon disulphide and alcohol, and has a specific gravity of 2.99. Symbol, Br; atomic weight, 80. Great care should be exercised in handling bromine.

Compounds.—Of the compounds containing bromine, the bromides of potassium, sodium, ammonium and hydrogen are the most largely used. They will all be considered under the respective bases.

TESTS FOR BROMINE AND BROMIDES.—Chlorine-water will liberate bromine from its combinations. Its odor is very characteristic. With starch it gives a yellowish color (iodine gives blue). Bromides, like iodides, give with silver nitrate a yellowish-white precipitate, insoluble in nitric acid and sparingly in ammonia-water. Strong sulphuric acid added to any dry bromide liberates vapors of bromine.

Fluorine.

THIS element is of little or no importance in medicine and pharmacy. It occurs chiefly as fluorspar (calcium fluoride). In its elementary state fluorine is scarcely known, because it so quickly and fiercely combines with all kinds of substances that chemists have had great difficulty in obtaining it in a free condition. Hydrofluoric acid attacks glass with such force that it is largely employed in the arts for etching on glass.

Sulphur.

Occurrence in Nature.—In Sicily and other parts of the world sulphur is found in the free or uncombined state, while as sulphides and sulphates it is very widely diffused. In its free condition it is found in the neighborhood of volcanoes, usually in opaque, crystalline masses, but sometimes in beautiful transparent yellow crystals. It is found as sulphides in combination with iron, copper, lead and mercury; as sulphates with calcium, barium, magnesium.

Preparation.—The *crude sulphur* is simply heated until it volatilizes, and the vapors cooled—that is, it is *sublimed*, which in a measure purifies it. Iron pyrites or iron sulphides yield a large quantity of sulphur when heated or sublimed.

Properties.—Sulphur when pure is a pale-yellow solid, permanent in the air, insoluble in water, soluble in disulphide of carbon, from which it may be crystallized.

When heated, sulphur melts and distils or sublimes, unaltered, if air is absent; in presence of air it takes up oxygen and becomes sulphur dioxide gas, SO_2 , just as it does when it burns. There are two forms of crystals of sulphur—an octahedral (eight-sided), in which the native sulphur sometimes occurs, and the monoclinic, obtained when a mass of sulphur is melted and, after partial cooling, the crust on the surface is broken and the fluid portion poured out. The crystals adhere to the inner portion and protrude into the space left by the portion which had not yet cooled and was poured out. The former is sometimes spoken of as the octahedral variety, and the latter as the prismatic variety. The former has a specific gravity of 2.04, while the prismatic weighs only 1.96 times as much as

water. Sulphur melts at 218° F., and boils or volatilizes at 752° F. It is interesting to notice that when sulphur is heated to the melting-point it forms a thin amber-colored liquid, which, when further heated, begins to thicken and becomes of a deeper color. At 430° or 450° F. it becomes very thick and tenacious, so much so that the vessel containing it may be inverted without losing its contents. If this now be poured into water and quickly cooled it becomes a very soft and flexible solid, which gradually hardens and becomes brittle. This condition may be looked upon as an amorphous one.

From the last-mentioned temperature up to its boiling-point, 752° F., sulphur again begins to become thin and limpid.

Sulphur, in its chemical relation, bears some resemblance to oxygen; many of the oxides have corresponding sulphides; for instance:

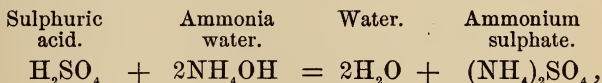
HgO, Mercuric oxide.	HgS, Mercuric sulphide.
CaO, Calcic oxide	CaS, Calcic sulphide.
FeO, Ferrous oxide.	FeS, Ferrous sulphide.
Fe ₂ O ₃ , Ferric oxide.	Fe ₂ S ₃ , Ferric sulphide.
PbO, Lead oxide.	PbS, Lead sulphide.

Official Sulphurs.—The pharmacopœia recognizes the *sublimed*, *washed* and *precipitated* sulphur, the ointment and the sulphur iodide.

SUBLIMED SULPHUR.—Vapors of sulphur when conducted into a cooled chamber become condensed in the form of a crystalline powder, which deposits upon the bottom and walls of the chamber, constituting the variety of sulphur known as the sublimed.

FLOWERS OF SULPHUR, as it is frequently called, usually

has an acid reaction, due to the absorption of oxygen and water, forming sulphurous and sometimes even sulphuric acid. This is not a very pure product; it may contain other impurities besides the acid, sometimes even traces of arsenic, especially if made from the iron pyrites. To insure purity the sublimed sulphur is washed with ammonia-water, yielding a product official under the name *Washed Sulphur*. The ammonia-water serves two purposes—it neutralizes the acid, forming sulphite and sulphate of ammonium:

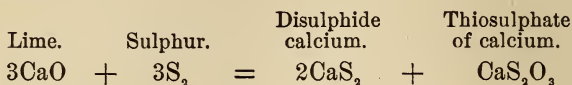


and dissolves out any arsenic that may be present. Subsequent washing with much pure water insures the total removal of the salts thus formed. The U. S. P. gives a test to ascertain the absence of arsenic. If the sulphur is digested with ammonia-water and filtered, the filtrate, if arsenic is present, turns cloudy when supersaturated with hydrochloric acid. This depends on the fact that arsenic is soluble in the alkali, but insoluble in the acid. The cloudiness would be due, therefore, to the precipitation in the acid solution of the arsenous sulphide. Or if the above filtrate be treated with sulphuretted hydrogen, made by acting on sulphide of iron in a test-tube with sulphuric acid, $\text{FeS} + \text{H}_2\text{SO}_4 = \text{FeSO}_4 + \text{H}_2\text{S}$, no precipitate should be produced, showing the absence of arsenous oxide.

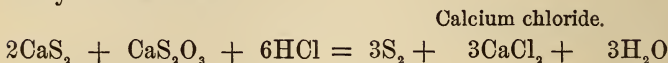
This form of sulphur is preferred by some for internal administration.

THE PRECIPITATED SULPHUR is much lighter and more easily suspended in liquids, and is therefore preferable to

the other forms of sulphur. It is made by boiling together the sublimed sulphur and lime, filtering the solution and adding hydrochloric acid:



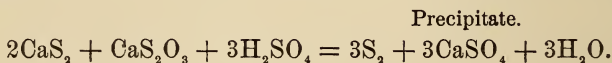
The solution containing the above calcium salts, when treated with hydrochloric acid, separates the sulphur in a finely-divided state:



The precipitate is thoroughly washed with much pure water.

The test applicable to the washed sulphur may be applied to this form of sulphur.

If sulphuric acid be employed in place of hydrochloric, the precipitated sulphur is contaminated with calcium sulphate, which is formed in the operation, and which, because it is insoluble, cannot, like the chloride of calcium, be washed out:



This kind of precipitated sulphur is known as *milk* or *lac sulphur*, and should not be mistaken for the official form.

IODIDE OF SULPHUR.—By rubbing together very thoroughly washed sulphur and iodine, and heating in a flask, carefully at first, not to dissipate the iodine; increasing the heat, when combination has taken place, to liquefaction; allowing to cool, breaking the flask and reducing the fused mass to small pieces, the iodide of sulphur is obtained. It is in the form of a grayish-black solid, and when exposed

to the air gradually loses its iodine, the latter being very loosely combined with the sulphur. Its use is chiefly externally, as ointment in skin diseases.

SULPHUR OINTMENT.—Made by rubbing sublimed sulphur with benzoinated lard. This is a convenient form for the administration of sulphur when the small proportion of acid present in the sulphur is of no consequence.

ALKALINE SULPHUR OINTMENT.—This was introduced into the 1880 pharmacopœia, and is made with *washed* sulphur potassium carbonate (the latter giving it its alkalinity, and hence its name) and benzoinated lard.

With *hydrogen* sulphur forms a gas known as sulphuretted hydrogen, or hydrosulphuric acid gas, which is of considerable importance in that it is one of the important reagents in pharmaceutical testing. Its composition is H_2S , one of those combinations in which sulphur has only two bonds, and which are known as *sulphides*, exemplified in the sulphide of iron, FeS ; sulphide of lead, PbS , etc. H_2S is an acid corresponding to those acids formed by the union of any of the halogens with hydrogen—that is, it is one of those acids containing no oxygen, and which belong to the hydrochloric-acid type, HCl . It may be appropriately mentioned here that in

The Writing of Salts

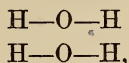
Or acids—that is, expressing their composition by means of symbols—among others, two types, the HCl type and the H_2O or water type, are employed.

For example, we may make, theoretically, a number of acids from hydrochloric acid by replacing its chlorine with some other one element: HCl , with the chlorine replaced by bromine, gives HBr , hydrobromic acid; in like manner,

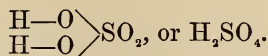
HI, hydriodic acid, HFl, hydrofluoric acid, etc., may be obtained. It will be noticed that the elements in the above acids all have one bond; an element replacing another must have an equal number of bonds in these cases. If the chlorine in the hydrochloric acid should be replaced by sulphur it would be necessary to employ at least two molecules of the acid, because sulphur, having two bonds, would replace two atoms of any element having only one bond; 2HCl or H_2Cl_2 with the Cl_2 replaced by one atom of sulphur would give H_2S . All acids, therefore, containing only two kinds of atoms may be said to be derived, theoretically, from one or more molecules of hydrochloric acid in which the chlorine is replaced by an equivalent of some other element. The hydrogen must not be replaced, because it forms the base of all acids, and when it is removed the combination is no longer an acid, as we shall see presently.

The other type, the water type, differs from the former in having *one* of the hydrogen atoms in water replaced usually by *two* or more other elements, one of which must be oxygen. Water is H_2O , or $\text{H}-\text{O}-\text{H}$; the right-hand hydrogen is the one that may be replaced. Nitric acid, HNO_3 , may be looked upon as $\text{H}-\text{O}-\text{H}$ in which the one H is replaced by NO_2 , thus, $\text{H}-\text{O}-\text{NO}_2$. The group of elements replacing this *one* hydrogen must have, *together*, *one* free bond; in this case the nitrogen has five bonds, and two atoms of oxygen together four bonds, which unite with four of the nitrogen, leaving one bond still free. NO_2 , therefore, is the equivalent in combining power of a single atom having one bond, and can replace the one atom of hydrogen in water, just as bromine can replace the chlorine in hydrochloric acid.

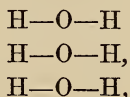
Sulphuric acid, H_2SO_4 , because it has two hydrogen atoms, is derived from two molecules of water:



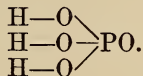
in which two atoms of hydrogen, one from each molecule of water, are replaced by SO_2 , sulphur dioxide:



Phosphoric acid is derived from three molecules of water,



in which the three atoms of hydrogen are replaced by PO :



In the case of sulphuric acid the S has six bonds, four uniting with the four of the two atoms of oxygen and the other two replacing the two hydrogens. Together, therefore, SO_2 has two free bonds. For a similar reason PO has three bonds, two of the five which phosphorus has uniting with the two bonds of the one oxygen, and the other three replacing the three hydrogen atoms in the three molecules of water. All acids may be thus classed in one of these two types, hydrochloric acid being the type for all acids containing no oxygen, while those containing oxygen are relegated to the water type. The student is earnestly advised to get the formulas of a number of acids—inorganic acids—from any text-book, and classify them into

the two types mentioned. There are other types to which reference may be had later on.

Now a step further. *Salts*, in a general sense, may be said to be acids in which the *hydrogen* has been replaced by a metal. Sulphate of magnesium, MgSO_4 , is a *salt* derived from sulphuric acid, H_2SO_4 , by replacing the hydrogen with magnesium. If the H in HNO_3 be replaced by sodium, Na, NaNO_3 , sodium nitrate, results; in a like manner, NaCl , sodium chloride, common table-salt, is HCl with the H removed and supplanted by Na, sodium or natrium.

In the case of monad metals, there must be as many atoms as there are hydrogen atoms in the acid:

HCl with Na yields NaCl , Sodium chloride.

HNO_3 with Na yields NaNO_3 , Sodium nitrate.

H_2SO_4 with Na yields Na_2SO_4 , Sodium sulphate.

H_3PO_4 with Na yields Na_3PO_4 , Sodium phosphate.

When the metal has more than one bond an equivalent of the acid must be employed:

2HCl with Ca yields CaCl_2 , Calcium chloride.

2HNO_3 with Ca yields $\text{Ca}(\text{NO}_3)_2$, Calcium nitrate.

H_2SO_4 with Ca yields CaSO_4 , Calcium sulphate.

$2\text{H}_3\text{PO}_4$ with 3Ca yields $\text{Ca}_3(\text{PO}_4)_2$, Calcium phosphate.

To make the salt calcium chloride we must remove the hydrogen from two molecules of hydrochloric acid, because Ca is a dyad; for the same reason we must employ two molecules of HNO_3 to make calcium nitrate. To make the calcium sulphate only one H_2SO_4 is necessary, because that contains the necessary H_2 . In trying to thus write the formula for phosphate of calcium the student may be puzzled a little at first, but he will have no trouble

if he will only remember that the base in the salt must have as many bonds as there are bonds in the base of the acid, and that he may employ more than one molecule of the acid if necessary. Ca has two bonds, but there are three hydrogens to be replaced. The easiest thing to do is to find the lowest multiple of the two numbers, in this case 2 and 3, which is 6. Six, then, is the number of hydrogens that there must be in the acid, hence we take the acid twice, $2\text{H}_3\text{PO}_4$. The base of the salt, then, also must contain six bonds; and 3Ca have six bonds, hence the formula of the salt will be $\text{Ca}_3(\text{PO}_4)_2$. In thus producing salts from acids there is merely an interchange of bases, of course governed by the number of bonds of the respective bases. It may be well to state here that a salt is by some looked upon as consisting of two parts, a *base* and a *radical*, or, as some chemists prefer, of a basylous and an acidulous radical. In potassium nitrate, KNO_3 , the K is the *base* or basylous radical, while the NO_3 is the *radical*, or acidulous radical. Acids are salts, but the term acid is specific, and means a salt containing hydrogen as its base.

The salts thus far spoken of are all *neutral* salts—that is, they are such that are derived from acids by replacing *all* the hydrogen of the latter by a metal. Now, it is possible to remove only part of the hydrogen with a metal, as is the case with bicarbonate of sodium, NaHCO_3 , which is carbonic acid, H_2CO_3 , with half its H replaced by sodium. Cream of tartar, $\text{KHC}_4\text{H}_4\text{O}_6$, is obtained by removing one H from tartaric acid, $\text{H}_2\text{C}_4\text{H}_4\text{O}_6$, and putting one of potassium in its place. Such salts are distinguished from the neutral salts by special terms; they are *acid salts* or *bisalts*. We distinguish another kind of salt, the *double* salt, exemplified in Rochelle salts, the tartrate of potassium and

sodium. Tartaric acid, $\text{H}_2\text{C}_4\text{H}_4\text{O}_6$, containing two hydrogen atoms in the base, has one replaced by potassium and the other by sodium, $\text{KNaC}_4\text{H}_4\text{O}_6$. Here the base consists of two metals.

The student is recommended to read again, very carefully, the chapter on Chemical Notation and Nomenclature.

But to return to sulphur. As already said, in H_2S we have a valuable reagent. When it comes in contact with the solutions of a great many of the metals it gives its sulphur to them, forming characteristic precipitates. Thus lead with H_2S gives a very black precipitate, while zinc yields a white one and arsenic a yellow one. These precipitates are *sulphides*—that is, the direct union of a metal with sulphur alone. H_2S may be made by acting on sulphide of iron with sulphuric acid: $\text{FeS} + \text{H}_2\text{SO}_4 = \text{FeSO}_4 + \text{H}_2\text{S}$. It has a very disagreeable odor, is soluble in water and in water of ammonia. The latter solution is also an official test-solution, that called ammonium sulphide. Sulphuretted hydrogen is the chief constituent of sulphur-waters. The sediment found in these waters at times is sulphur, due to the decomposition of the gas, setting the sulphur free.

The other acids of sulphur will be considered in the chapters on acids.

Phosphorus.

Occurrence in Nature.—Phosphorus never occurs in a free state in Nature, but is most usually found in the form of phosphates of some of the metals, principally of calcium, sodium, potassium, iron and aluminum. The soil contains

a considerable quantity, and is fertile in proportion to the amount of it present. Soil, it should be understood, is the result of the disintegration and tumbling down of the ancient unstratified rock, and of some of modern origin, and these ancient rocks having contained phosphorus as phosphates, the latter is now present in soil as a consequence. And that this is so is another illustration of the sublime economy of nature, for phosphorus is an essential constituent of plant food. Most plants could not exist were it not for the presence of phosphorus in the soil—this is shown on fields in which there have been many successive crops without the aid of artificial fertilizers—each successive crop becomes less abundant, because each previous one has diminished the amount of phosphorus present. The phosphorus which thus passes into the organism of plants finds its way into the bodies of animals to which plants serve as food. The bones of animals are made up almost wholly of phosphate of calcium, thus indirectly derived from the soil through the intervention of the plant world. The salt plays an important part in the animal frame, the skeleton of man and animals being largely composed of it.

Source.—Phosphorus is chiefly obtained from bones, which contain 60 per cent of phosphate of calcium, or tricalcic phosphate, $\text{Ca}_3(\text{PO}_4)_2$. Formerly it was obtained from urine. Brandt, of Hamburg, a chemist, discovered it in this excretion in 1669, and for a long time that was its only source. On account of the small yield it was very scarce and expensive at first.

Preparation.—Nowadays phosphorus is prepared from bones, which are first calcined and then reduced to a fine powder. Calcination removes the fat and all other organic

matter which heat will destroy, leaving as residue a salt which is mainly calcium phosphate. The powdered calcined bones are then mixed with sulphuric acid diluted with water, and the mixture allowed to stand, to permit the insoluble sulphate of calcium which is formed to deposit. The tricalcic phosphate (having three atoms of calcium) $\text{Ca}_3(\text{PO}_4)_2$, becomes converted into a monocalcic (mono = one) phosphate, which is soluble and remains in solution: $\text{Ca}_3(\text{PO}_4)_2 + 2\text{H}_2\text{SO}_4 = \text{CaH}_4(\text{PO}_4)_2 + 2\text{CaSO}_4$.

After standing for several hours the supernatant liquid is poured off from the deposited calcium sulphate, filtered and evaporated to a syrupy consistency, mixed with powdered charcoal, and the whole thoroughly dried in a vessel exposed to a high temperature. When quite dry the residue is introduced into a stoneware retort and carefully heated again; vapors of phosphorus are distilled and carefully collected under water



The $\text{Ca}_3(\text{PO}_4)_2$ remains in the retort together with some charcoal; the mixture is incinerated, thereby removing the charcoal and fitting the bone phosphate for treatment with sulphuric acid, to convert it again into the soluble phosphate. This method is the one usually employed, and by it 100 pounds of fresh bones yield from 6 to 7 pounds of phosphorus.

Another method resembling the above is coming into very frequent use. Instead of mixing the soluble phosphate with charcoal alone, sand is added and the process conducted as above. (Sand is mainly oxide of silica): $2\text{CaH}_4(\text{PO}_4)_2 + 2\text{SiO}_2 + 10\text{C} = 2\text{CaSiO}_3 + 10\text{CO} + 4\text{P} + 4\text{H}_2\text{O}$. The soluble phosphate, $\text{CaH}_4(\text{PO}_4)_2$, is probably first converted

during the process of heating into the metaphosphate: $\text{CaH}_4(\text{PO}_4)_2 + \text{heat} = \text{Ca}(\text{PO}_3)_2$, metaphosphate of calcium, and $2\text{H}_2\text{O}$, water.

Extreme care must be exercised in manipulation in these processes; they should not be attempted by the student.

Properties.—Phosphorus is a soft, flexible solid, colorless, translucent when recently prepared, and of the appearance of imperfectly bleached wax. When exposed to light and time it gradually becomes less translucent, finally opaque, white, yellow and finally red. The latter conditions are allotropic ones, which will be alluded to again. Phosphorus has a specific gravity of 1.77 to 1.83, melts at 44°C ., and boils at 280°C . When it is slowly cooled it sometimes crystallizes in well-formed dodecahedrons. It is insoluble in water, soluble in oils, ether, chloroform and very soluble in disulphide of carbon. In air it very quickly oxidizes and takes fire spontaneously. Its name signifies "bearer of light," and was given it on account of its property of being luminous in the dark. "If touched it took fire and burned furiously, exhaling a dense white cloud, which gathered like fleeces of snow, but, unlike snow, hissing like a red-hot iron when touched with water, or, if brought into contact with the skin, blistering it like living fire." In this way did Brandt refer to its properties after he discovered it. They spoke of it at the time as "the son of Satan." It is exceedingly inflammable, taking fire from the heat of the hand or from a blow or hard rub. Sunshine ignites it immediately. This is due to its great affinity for oxygen, with which it combines very energetically, and, in consequence of which property, it has to be kept under water, upon which it has no effect. A stick of

phosphorus, when exposed to the air, always emits a whitish smoke or cloud, due to a slow combustion which the phosphorus undergoes with the oxygen of the air, and upon which one of the methods for the analysis of the air depends. The white fumes which it emits when burning are P_2O_5 , phosphoric oxide, which will be referred to again. The burns caused by phosphorus are exceedingly painful and produce sores which are difficult to heal; the greatest care must be taken in handling it.

ALLOTROPIC CONDITIONS.—Among the remarkable properties of this singular substance is the diversity of its allotropic conditions, of which it assumes five. The ordinary translucent vitreous phosphorus when exposed to light under water changes to the *first* allotropic variety, which is white, opaque and less fusible. The *second* variety is the crystalline, obtained by evaporating some of its solutions; the *third*, a black, opaque variety, produced by the sudden cooling of melted phosphorus: a *fourth* allotropic condition is obtained by suddenly cooling phosphorus when it is near its boiling-point, when is yielded a soft, elastic substance, analogous to viscous sulphur, described when treating of sulphur; the *fifth*, called *red phosphorus*, is obtained by exposing phosphorus to the sun under water, or by exposing it to the sun between plates of glass. The latter amorphous variety may also be obtained by heating common phosphorus in carbonic acid gas, or in any atmosphere that will not act on it chemically. It may be looked upon as the passive condition of phosphorus, as it is devoid of many of the energetic properties of the ordinary or active variety. It is red in color, heavier than the former, is not luminous, nor does it melt at the boiling-point of water. It exhales no vapors or odor, oxidizes but very little in the

a.r., does not change oxygen into ozone, is chemically indifferent toward other elements, may be handled with impunity or carried exposed in the pocket, and is not poisonous when administered in doses a hundred times larger than would be fatal in the common form (G. Wilson). It is peculiar also in that it changes into the active form and bursts into flame at 500° F.

POISONOUS PROPERTIES OF PHOSPHORUS.—This element is relegated to the list of active poisons, and the student cannot be often enough impressed with its toxic properties. Taken internally in larger than therapeutic doses it acts as a violent poison, being all the more toxic because there is no specific antidote for it. Two kinds of phosphorus poisoning are usually distinguished—the acute form and the chronic. The former results upon ingestion of an excessive amount, while the chronic form affects especially workmen employed in the manufacture of matches, an industry which is very extensive. Oil of turpentine has been recommended as an antidote, but some authors oppose its use, on the ground that it dissolves the phosphorus, rendering it even more active. The stomach-pump and emetics and, later, cathartics should be employed to eliminate as speedily as possible the phosphorus in a mechanical way. It must be borne in mind that this poison is soluble in oils or fats, and hence these should not be given, not even milk, since that contains considerable fat.

Experiments.—The student is fully aware by this time of the danger connected with the handling of phosphorus, but yet he should have enough confidence to undertake a few experiments which will make him learn to exercise care in the manipulation of a dangerous substance.

Put into a flask, having a well-fitting cork, a small piece of phosphorus by means of small pincettes, add two or three drams of ether and allow to stand for a few days, shaking occasionally. The ether will dissolve most of the phosphorus. Several drops of this solution poured upon the hands and these quickly rubbed together will cause the ether to evaporate and the phosphorus to remain upon the hands in a state of minute division. The phosphorus soon begins to become luminous, causing the hands to shine in the dark, the luminosity becoming more intense upon rubbing the hands, as a fresh surface of phosphorus is thus continually presented to the oxygen of the air. Heat is produced, but it is too feeble to occasion ignition; it is slow combustion. The phosphorus unites chemically with the oxygen of the air, producing one of the oxides of phosphorus, which in turn combines with the water in the air to produce phosphoric acid. If a lump of sugar be moistened with the ethereal solution of phosphorus and thrown into hot water, the heat of the latter causes the ether to volatilize, and the phosphorus rises to the top and there inflames spontaneously. In this case the combustion is brisk and complete, and the phosphorus takes up a larger amount of oxygen than above, forming the phosphoric oxide, P_2O_5 , and this with the moisture of the air, or with water, produces phosphoric acid. In the preceding experiment the oxide produced was the phosphorous oxide, P_2O_3 , which is always produced when phosphorus combines slowly and incompletely with oxygen, while, when the oxidation is complete—that is, when it is accompanied by flame—the higher phosphoric oxide, P_2O_5 , is always formed.

Pour some of the ethereal solution of phosphorus upon

fine blotting-paper, and as soon as the ether will have evaporated the paper will ignite spontaneously.

If a little finely-powdered charcoal or soot is sprinkled over a small piece of phosphorus the latter melts after a while and spontaneously inflames. The charcoal causes this combustion; because of its porosity it absorbs oxygen from the air, imparts it to the phosphorus, and being also a non-conductor of heat prevents the cooling of it.

A piece of phosphorus in a wine-glass will melt if hot water is poured upon it, but will not ignite because of the absence of air. If air be blown very carefully through a long tube upon the phosphorus a combustion will take place which is especially visible in the dark. An oxide of phosphorus is formed, but a much lower one than in any of the experiments, the phosphorus being in excess of the oxygen, in this instance, P_2O . The three oxides mentioned thus far, the P_2O , P_2O_3 and P_2O_5 , will receive more attention later on, under Phosphoric Acid.

The P_2O may also be obtained by gently heating a small piece of phosphorus placed in the middle of a glass tube about a foot or 15 inches long. When ignition commences the heat should be withdrawn. The combustion is feeble and imperfect while the tube is held horizontally, because the heavy smoke, consisting of phosphoric and phosphorous oxides, passes off slowly, allowing the admission of only a small quantity of air. As soon as the tube is inclined the combustion becomes more vivid, and when the tube is held perpendicularly it becomes perfect, because of the draught of air through the tube. In this wise phosphorus may be oxidized to any of the three degrees.

Uses of Phosphorus.—In the manufacture of matches we have one of the chief uses for phosphorus. Vast quantities

are consumed in this way. In making the matches the wood is cut by machinery, the ends first tipped with sulphur, which is in a molten condition for the purpose, and then with an emulsion of phosphorus in ordinary glue, to which is added a little nitrate of potassium, oxide of manganese or chlorate of potassium. The latter bodies are all rich in oxygen. In igniting a match, the friction-heat produced by striking it against a surface is sufficiently high to cause the phosphorus to kindle and unite with the oxygen furnished by the nitrate or chlorate of potassium. The heat and flame produced by this chemical union between the phosphorus and oxygen are sufficient to cause the sulphur to burn, which in turn ignites the wood. Not only is the manufacture of matches dangerous from the explosive nature of the materials used, but also because of the corrosive phosphoric vapors, which produce at times among the laborers a distressing disease known as caries of the lower-jaw. Improvements in the methods employed have lessened the danger of this dread disease.

In treating of the occurrence in nature of phosphorus, we learned how essential it was to plant and animal life, how it came to be in the soil, and how it found its way through the vegetable world into the animal, where its function is most important in that it builds up the bony structure. It is also an important constituent of the brain and nervous system. After continued brain exercise there is always an increase in the proportion of phosphoric products in the liquid excretions.

Phosphorus in Pharmacy and Medicine.—Phosphorus is official in four forms: as phosphorus in its unaltered state, as a solution in the form of Phosphorated Oil, in a minutely-divided form in the Pills of Phosphorus, and in a state of

combination with oxygen and hydrogen in Phosphoric Acid. There are other compounds besides these in which phosphorus occurs, but they are not made directly from phosphorus.

The Pharmacopœia directs phosphorus to be kept under water, in a secure and moderately cool place, protected from light.

When phosphorus is intended for medicinal use it should be absolutely pure. The commercial article sometimes contains a little arsenic or sulphur, or both, which render it totally unfit for the purposes of medicine, and the presence of which should be strenuously guarded against. To detect arsenic, a portion of phosphorus should be dissolved in nitric acid, water added, and the solution evaporated until all superfluous nitric acid is dispelled. Before the solution (which should be somewhat diluted) has thoroughly cooled a stream of hydrosulphuric acid gas, H_2S , should be passed through it for half or three quarters of an hour. If arsenic is present a lemon-yellow deposit will be formed within twenty-four hours, abundant in proportion to the quantity of arsenic. If no deposit forms within that time, it may be assumed that there was no arsenic present. The Pharmacopœia allows a slight quantity of sulphur to be present in phosphorus. If to the above solution a few drops of test-solution of chloride of barium be added, not more than a slight opalescence should make its appearance. The opalescence is due to the presence of sulphuric acid, which is formed by the action of the nitric acid upon traces of sulphur which are present.

The dose of phosphorus is from $\frac{1}{100}$ to $\frac{1}{30}$ grain. In large doses it acts as a poison. The medicinal value of phosphorus as a nervous stimulant seems only to be

obtained by administering it in a free state. The oxide of phosphorus, or phosphoric acid, has a different action, and hence the necessity of preventing the oxidation of the phosphorus. The pills and the phosphorated oil are intended for the administration of phosphorus in a free state.

PHOSPHORATED OIL is prepared by heating expressed oil of almond to about 250° C., allowing to cool, filtering, adding the phosphorus, heating again in a water-bath until the phosphorus dissolves, allowing to cool and adding ether. Phosphorated oil, when fresh, should be clear and colorless, or only slightly colored, phosphorescent in the dark, and have the taste and odor of phosphorus. The oil is first heated to expel air or water, which, if present, would oxidize the phosphorus. The ether, besides serving to preserve the oil, also renders its taste less disagreeable. A good way of administering phosphorated oil is in milk or in the official emulsion of almonds; the dose is from 1 to 5 minims, the oil containing 1 per cent phosphorus.

PHOSPHORUS PILLS contain $\frac{1}{100}$ grain each, and are made by dissolving the phosphorus in chloroform, adding this chloroformic solution to a mixture of marshmallow and acacia in a mortar, and quickly making a mass with glycerin and water. To prevent the oxidizing action of the air they are directed to be coated with balsam tolu which has been dissolved in stronger ether. The pills are shaken with a sufficient quantity of the tolu-ether solution and put away on plates to dry. The balsam forms a coating which is impervious to air or moisture. The pills should be kept in small, well-stoppered bottles. In making the solution

care should be taken not to use heat, as ether is very inflammable and volatile.

PHOSPHORIC ACID also is made directly from phosphorus, but that will be referred to in the chapter on acids.

Boron.

BORON is of some importance in pharmacy, being the chief chemical constituent of the important and much used borax and boric acid, and of all borates.

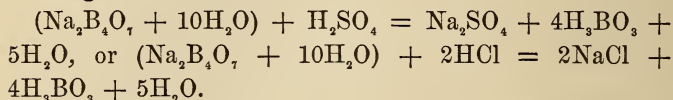
Occurrence in Nature.—Boron is never found in the free state in nature, but most abundantly in boric acid and in sodium borate. The boric acid occurs in solution in the hot lagoons of Tuscany, while borax forms a crystalline deposit in many of the lakes of California and in those of Thibet and Persia. The borax obtained from the latter is imported in a crude state under the name of *tinca*, which is crystallized and purified here, though the main supply of our market is obtained at present from the lakes of California. A native borate of calcium is found in Southern Peru; this and the native boric acid from Tuscany furnish a large proportion of the borax in the market. Besides these sources, *borax springs* furnish considerable borax. It is a peculiar fact that borax springs always contain ammonia.

Preparation.—Boron is obtained by heating one of its compounds, boric acid, preferably, with magnesium in a suitable tube, and by continued washing in dilute hydrochloric acid, washing out the borate and boride of magnesium which are formed.

Boric Acid.—There are many compounds containing boron, but we confine ourselves to the two of importance in pharmacy, namely *boric acid* and *borax*. The lagoons

of Tuscany formerly furnished the greater part of the boric acid in commerce, but at present the most of it is obtained from the borax found in the lakes of California by decomposing it with hydrochloric acid.

Borax is a baborate of sodium, $\text{Na}_2\text{B}_4\text{O}_7 \cdot 10\text{H}_2\text{O}$, which when treated with either hydrochloric or sulphuric acid, while in hot solution, has its sodium displaced by the acid, the hydrogen of the acid combining with the boric radical, forming boric acid, thus :



In the first instance sulphate of sodium is formed as a by-product; in the second, a chloride. Both these salts are very soluble and remain in solution, while the much less soluble boric acid separates upon cooling, and is purified by redissolving in pure water and again crystallizing.

As thus obtained, boric acid is in transparent six-sided plates or scales, soluble in 25 parts of water; much more soluble in hot water. It is a weak acid compared with the other mineral acids, and turns blue litmus paper not red, like other acids, but only a wine-red color; while turmeric paper (porous paper soaked in solution of turmeric) becomes brown in its contact with it. Alkalies have nearly the same effect on turmeric paper, but the brown color disappears on the addition of stronger acids. Turmeric paper turned brown with boric acid does not lose its color with acids, but turns bluish-black with ammonia-water.

Boric acid, added to alcohol, imparts to the flame of the latter a green color. This is a characteristic property of

boric acid. The crystallized acid contains water of crystallization— $\text{H}_3\text{BO}_3 + \text{H}_2\text{O}$ —and when heated melts in this water, which volatilizes, leaving a glassy bead of anhydrous boric acid, which is much utilized in blowpipe analysis. A piece of platinum wire, with a loop on the end, if wetted and dipped into boric acid, affords a convenient means of exposing boric acid to the blowpipe flame. If a stream of air is directed through the blowpipe into the flame of a spirit lamp, or Bunsen-burner, and the wire with the acid brought near to the flame, the acid will first melt and swell in its water of crystallization, and finally will be converted into a transparent glassy bead. This bead consists of the oxide of boron, probably B_2O_3 , which, in its molten condition, has the property of dissolving many metallic oxides with very marked and characteristic colors. If this glassy bead be moistened and dipped into chalk or litharge, or into almost any oxide, and again heated to melting, these substances will unite most intimately with the boric oxide, and be dissolved by it, and assume a glassy, or, better, *vitri-fied* appearance. Most of the combinations of boric acid with oxides or bases become vitreous on heating—that is, they melt together, forming a white or colored glass. Chromium salts thus give a green color, manganese an amethyst, and cobalt a blue, etc.

The official boric acid is only one of the many acids of boron, which will receive mention under BORAX.

THE BLOWPIPE is a very useful instrument for melting, oxidizing, reducing or volatilizing substances in small quantities. It is made in various shapes, but consists essentially of a metal tube bent a short distance from one end and drawn to a point. The proper use of the instru-

ment requires a little practice, which can be soon acquired. The blast should be steady and continuous, and should be produced principally by the muscles of the cheek, the respiration not to interfere much. A double combustion takes place in the blowpipe flame, a smaller *interior* cone of a blue color and a larger *exterior* cone of a yellowish appearance. The latter is called the *oxidizing* flame, the former, the *reducing*. If a substance is to be *reduced*—that is, deprived of its oxygen—it is held at the point of the inner flame, where it meets with soot or carbon, which combines with the oxygen of the substance and escapes as carbonic acid, leaving the substance reduced, or devoid of oxygen. If, on the other hand, a body is to be *oxidized*—that is, united with oxygen, then it is held at the point of the outer flame, where the oxygen of the air can have free access to it. It may be good practice for the student to convert a piece of lead, placed upon charcoal, into an oxide, by exposing it to the outer flame, and then reducing it, by exposing the oxide thus formed, to its original metallic state.

Tests for Purity of Boric Acid.—The Pharmacopœia requires boric acid, or, as it was formerly called, *boracic* acid, to be free from sodium and from sulphuric and hydrochloric acids, which may be present as a result of incomplete washing after the acid is separated from its solution with either the chloride or sulphate of sodium. Lead, copper and calcium should also be absent. Barium chloride solution will give a white precipitate with sulphates, insoluble in any acid. Solution of silver nitrate yields a curdy white precipitate with chlorides, soluble in ammonia-water and precipitated upon addition of nitric acid. Calcium may be detected by the addition of a solution of ammonium

oxalate, which produces a white precipitate with soluble calcium salts.

Lead, copper and iron give a black color and precipitate, with ammonium sulphide.

The above tests are all characteristic, and not only applicable in these cases, but in many others where the same substances are sought. It is well, therefore, that the student commit them to memory.

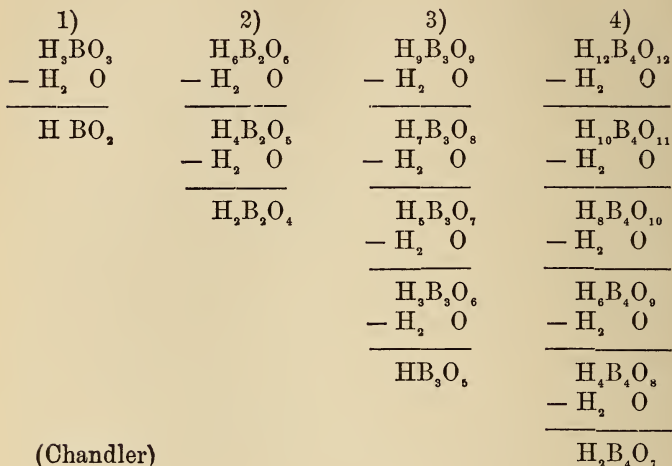
Uses in Pharmacy and Medicine.—Boric acid is required to be in a very fine impalpable powder for most medical uses. It is a good antiseptic, and is extensively employed in antiseptic surgery. Borated cotton, lint or gauzes and bandages are much used as antiseptic dressings to wounds. It is an excellent means of preventing fermentation, and is for that reason added to many liquids, to which it communicates but little taste. In form of ointment or dusting-powder it is much favored for external application.

Borax.

Borax is a sodium salt, usually termed in chemical language the borate or baborate of sodium. Its composition is expressed by the formula $\text{Na}_2\text{B}_4\text{O}_7 \cdot 10\text{H}_2\text{O}$, which shows it to be composed largely of water of crystallization. It contains four atoms of boron, and may be looked upon as a salt of one of the many boric acids before referred to.

As an illustration of the many acids of boron the following may serve: Beginning with one molecule of orthoboric acid, H_3BO_3 , we may obtain another by abstracting one molecule of water. Beginning with successively two, three and four molecules of the orthoboric acid (ortho means "straight;" that is, "the" acid) and abstracting molecules

of water as long as possible, we get a large variety of boron acids, viz.:



The last acid, $\text{H}_2\text{B}_4\text{O}_7$, may be looked upon as the one furnishing borax or borate of sodium by having its hydrogen replaced by as many sodium atoms, $\text{Na}_2\text{B}_4\text{O}_7$. The $10\text{H}_2\text{O}$ are taken up during crystallization. $\text{H}_2\text{B}_4\text{O}_7$ is called the *pyroboric acid*, but the term is applicable to any of the other acids obtained by abstracting water from the ortho acid by heating, pyro meaning "fire;" that is, in this case, an acid obtained by *heating* another.

Pyrophosphoric acid is another example of a pyro acid.

Borax, as already mentioned, is obtained principally from the borax lakes in California, where it occurs as a crystalline deposit in the mud at the bottom of the lakes. The large crystals are washed with water, dissolved and crystallized, thereby yielding an almost absolutely pure

salt, which, when powdered, constitutes the powdered borax of commerce.

Borax is also made from the boric acid obtained from Tuscany by fusing it with dried sodium carbonate and crystallizing the hot solution of the residue.

The official borax is in colorless, shining, monoclinic prisms, somewhat efflorescent in dry air. It has an alkaline reaction, is soluble in 16 parts of cold and very soluble in boiling water, insoluble in alcohol. It is very soluble in warm glycerin and honey.

Tests for the Purity of Borax.—Some of the possible impurities in borax are metals, alkaline earths, carbonates, sulphates and chlorides.

Sulphuretted hydrogen proves the absence of metals if it does not affect a solution of borax.

The alkaline earths (Ca, Ba, Sr, etc.) are precipitated, if present, by a solution of sodium carbonate. Barium and calcium, if present, are in minute quantity and dissolved in the solution of borax. When sodium carbonate is added they are converted into the less soluble carbonates of barium and calcium, which precipitate.

An acid added to a solution of borax causes an effervescence if carbonate is present. (When acids come in contact with any carbonate this happens always: $\text{Na}_2\text{CO}_3 + 2\text{HCl} = 2\text{NaCl} + \text{H}_2\text{CO}_3$. The H_2CO_3 is carbonic acid, which always resolves itself into H_2O , water, and CO_2 , carbonic acid gas. The escape of the gas produces what is known as *effervescence*, by carrying with it, in its endeavor to get away, portions of the liquid, which when intermingled with the gas constitute *foam*.)

Sulphates and chlorides are detected, as in boric acid, with barium chloride and silver nitrate, respectively.

Borax enters into many mouth and tooth washes—the *honey and borax* used so extensively, and made by dissolving borax in honey, is an excellent form for the use of borax as a mild antacid, especially in the sprue or thrush of children. It is frequently used to whiten ointments; the yellow color which cold cream sometimes takes may be prevented by the addition of a little borax solution during its preparation.

Inorganic Acids.

At this stage of the course the student is earnestly advised to read over again very carefully all of the foregoing chapters, especially those on chemical philosophy, as he is about to enter upon the study of the *acids*, in which he will find the application of the principles of chemistry of constant necessity. The non-metallic elements only have thus far attracted his attention; these mastered, the student is ready to go deeper into the details of

The Study of the Acids and Salts.

When treating of oxygen, we learned that an *oxide* is the union of oxygen with another element. We will go further now and divide the oxides into three principal groups or classes. The first class includes all those which resemble, chemically, the oxides of sodium, potassium, barium or ammonium, and others, and are called *basic oxides* or *alkali oxides*. The second division comprises those oxides whose properties are in opposition to those of the first, and which are represented by the oxides of sulphur, phosphorus and carbon—these are the *acid oxides*. The third class is represented by a single compound—water, H_2O —the oxide of hydrogen. When the first two oxides combine—i.e., the basic with the acid—a *salt* is produced: Oxide of calcium, CaO , uniting with the acid oxide of carbon, CO_2 , furnishes a salt—namely, $CaCO_3$. In like manner, for illustration, the following basic oxides uniting with acid oxides form the salts mentioned :

	Basic oxides.	Acid oxides.	Salts.	
Potassium oxide,	$\left\{ \begin{array}{l} \text{K}_2\text{O} + \text{SO}_2 = \text{K}_2\text{SO}_3, \\ \text{Na}_2\text{O} + \text{CO}_2 = \text{Na}_2\text{CO}_3, \\ \text{BaO} + \text{SO}_2 = \text{BaSO}_3, \\ 3\text{K}_2\text{O} + \text{P}_2\text{O}_5 = 2\text{K}_3\text{PO}_4, \end{array} \right\}$	SO_2	$= \text{K}_2\text{SO}_3,$	$\left\{ \begin{array}{l} \text{Potassium} \\ \text{sulphite.} \end{array} \right\}$
Sodium oxide,		CO_2	$= \text{Na}_2\text{CO}_3,$	$\left\{ \begin{array}{l} \text{Sodium} \\ \text{carbonate.} \end{array} \right\}$
Barium oxide,		SO_2	$= \text{BaSO}_3,$	$\left\{ \begin{array}{l} \text{Barium} \\ \text{sulphite.} \end{array} \right\}$
Potassium oxide,		P_2O_5	$= 2\text{K}_3\text{PO}_4,$	$\left\{ \begin{array}{l} \text{Potassium} \\ \text{phosphate.} \end{array} \right\}$

This is one way of producing a salt.

When the acid oxides come in contact with water (the third division of the oxides) union takes place, with the formation of *acids*:

	Acid oxides.	Water.	Acids.	
Sulphurous oxide,	$\left\{ \begin{array}{l} \text{SO}_2 + \text{H}_2\text{O} = \text{H}_2\text{SO}_3, \\ \text{CO}_2 + \text{H}_2\text{O} = \text{H}_2\text{CO}_3, \\ \text{P}_2\text{O}_5 + 3\text{H}_2\text{O} = 2\text{H}_3\text{PO}_4, \\ \text{N}_2\text{O}_5 + \text{H}_2\text{O} = 2\text{HNO}_3, \end{array} \right\}$	H_2O	$= \text{H}_2\text{SO}_3,$	$\left\{ \begin{array}{l} \text{Sulphurous} \\ \text{acid.} \end{array} \right\}$
Carbonic oxide,		H_2O	$= \text{H}_2\text{CO}_3,$	$\left\{ \begin{array}{l} \text{Carbonic} \\ \text{acid.} \end{array} \right\}$
Phosphoric oxide,		$3\text{H}_2\text{O}$	$= 2\text{H}_3\text{PO}_4,$	$\left\{ \begin{array}{l} \text{Phosphoric} \\ \text{acid.} \end{array} \right\}$
Nitric oxide,		H_2O	$= 2\text{HNO}_3,$	$\left\{ \begin{array}{l} \text{Nitric} \\ \text{acid.} \end{array} \right\}$

Acids are of two principal kinds—the *hydracids*, those containing no oxygen, and the *oxyacids*, those containing oxygen and represented by the above. A third division is sometimes made, including those represented by arsenious acid, As_2O_3 . The latter had better be termed *anhydrides*, because they contain no hydrogen, and hydrogen is an essential constituent of all acids. The word acid is a misnomer chemically for all substances that do not contain replaceable hydrogen in their base.

We have learned that one way of producing salts is by uniting a basic with an acid oxide; another is by replacing

the hydrogen in acids with metals, thus (Free hydrogen is not the by-product in each case, but is assumed to be to illustrate the lesson in hand):

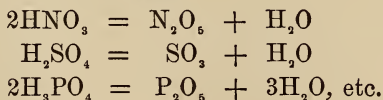
Acid.	Metal.	Salt.
Nitric, 2HNO_3 ,	$\} + \text{Sodium, Na}_2$,	$= \{ 2\text{NaNO}_3$, Nitrate sodium, + H_2 .
Hydrochloric, 2HCl ,	$\} + \text{Zinc, Zn}_2$,	$= \{ \text{ZnCl}_2$, Chloride zinc, + H_2 .
Sulphuric, $2\text{H}_2\text{SO}_4$,	$\} + \text{Iron, Fe}_2$,	$= \{ 2\text{FeSO}_4$, Sulphate iron, + 2H_2 .
Phosphoric, $2\text{H}_3\text{PO}_4$,	$\} + \text{Calcium, 3Ca}$,	$= \{ 2\text{Ca}_3(\text{PO}_4)_2$, Phosphate cal- cium, + 3H_2 .

And a third by replacing the hydrogen with basic oxides, thus:

Acid.	Basic oxide.	Salt.
2HNO_3	+ Sodium oxide, Na_2O ,	$= 2\text{NaNO}_3$ + H_2O .
2HCl	+ Zinc oxide, ZnO ,	$= \text{ZnCl}_2$ + H_2O .
H_2SO_4	+ Iron oxide, FeO ,	$= \text{FeSO}_4$ + H_2O .
$2\text{H}_3\text{PO}_4$	+ Calcium oxide, 3CaO ,	$= \text{Ca}_3(\text{PO}_4)_2$ + $3\text{H}_2\text{O}$.

The three ways of making salts, then, are: 1, by uniting basic and acid oxides; 2, by uniting metals with acids; and 3, by uniting basic oxides with acids.

In the last-mentioned method it will be noticed that the oxide of hydrogen—water—is liberated, and that only the acid oxides in the respective acids do the real work for the acids—i.e., oxyacids resolve themselves during the union into their acid oxides and water:



Acids are salts. They are a particular group, and the oxyacids always contain the elements of an acid oxide and hydrogen oxide. Many of them possess in a marked degree the properties to which the term "acid" is applied, such

as sour taste, solubility in water, corrosive action, power of reddening vegetable colors, and neutralizing alkalies or alkaline bases, basic oxides. The most characteristic property of the acids is their power of exchanging their hydrogen for metals or basic oxides to form salts—basic oxides are often spoken of as *bases*—as shown above. Thus when sulphuric acid, which contains hydrogen, sulphur and oxygen, comes in contact with a metal—zinc, for example—the hydrogen is replaced by the zinc, and that which was hydrogen sulphate becomes zinc sulphate, the hydrogen being liberated. The same will happen when a base comes in contact with an acid, only that the hydrogen which is set free unites with the oxygen combined with the metal, and forms water.

It must be borne in mind by the student that these changes all take place according to the laws of chemical combination. The student should also refresh his memory regarding the atomicity, number of bonds, etc., of all the elements.

Salts, and therefore acids, are usually spoken of as containing a *base* and a *radical*; in H_2SO_4 , for instance, the H_2 is the base and SO_4 the radical; in NaNO_3 the Na is the base and NO_3 the radical. (The term base should not be confounded with the term as used before.) It would be more correct to speak of an *acidulous* radical, and a *basylous* radical as constituting an acid—that is, it would be more rational, as an *acid* and a *base* when united form a salt, and both may be looked upon as radicals or roots.

In writing salts the base, as well as the radical, has a certain number of bonds, whether each consists of only one element or of more. In the following there is only one element in each, the base and the radical, and consequently

the base and radical have each as many bonds as their respective elements :

Base.	Radical.	
Sodium'	Chloride'	NaCl
Hydrogen'	Bromide'	HBr
Potassium'	Iodide'	KI
Calcium''	Chloride''	CaCl ₂

In the following there is one element in the base and more than one in the radical. The elements in the *radical* form a group which may be looked upon as behaving as a single element does and together having a certain number of free bonds.

	Base.	Radical.		
HNO ₃	H'	(NO ₃)'	nitric acid	} Inorganic.
H ₂ SO ₄	H ₂ ''	(SO ₄)''	sulphuric acid	
H ₃ PO ₄	H ₃ '''	(PO ₄)'''	phosphoric acid	
HC ₂ H ₃ O ₂	H''	(C ₂ H ₃ O ₂)''	acetic acid	} Organic.
H ₂ C ₄ H ₄ O ₆	H ₂ ''	(C ₄ H ₄ O ₆)''	tartaric acid	
H ₃ C ₆ H ₅ O ₇	H ₃ '''	(C ₆ H ₅ O ₇)'''	citric acid	

The (NO)' has one bond, the (SO₄)'' two, and the (PO₄)''' three, etc.; these radicals interchange in the different combinations without altering their relative proportion or position. The student should study the valence of the above-mentioned acid radicals and commit them to memory. In order to make salts from these or other acids it is only necessary to replace the hydrogen with an equivalent quantity of some metal, and to remember that the hydrogen alone can be replaced.

Bases, like radicals, may consist of more than one element :

	Base.	Radical.	
Sodium bicarbonate	(NaH)''	(CO ₃)''	NaHCO ₃
Potassium bitartate.....	(KH)''	(C ₄ H ₄ O ₆)''	KHC ₄ H ₄ O ₆
Sodio-potassic tartrate.....	(KNa)''	(C ₄ H ₄ O ₆)''	KNaC ₄ H ₄ O ₆
Acid calcium phosphate	(CaH)'''	(PO ₄)'''	CaHPO ₄

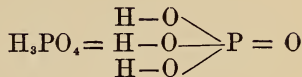
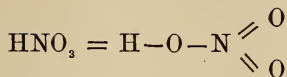
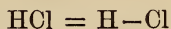
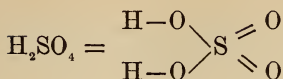
These salts are known as *acid* or *bi-salts* when they contain some unreplaced hydrogen, and as *double salts* when the hydrogen of a dibasic or tribasic acid is replaced by two metals, as in Rochelle salt, KNaC₄H₄O₆, etc. Acids having one replaceable hydrogen atom are called *mono-basic*, as HCl, HNO₃; when they contain two replaceable atoms, *dibasic*, as H₂SO₄; when three, *tribasic*, etc. The same terms apply to salts.

In writing the formulas for salts or acids, the base occupies the left and the radical the right. Thus, sulphate of sodium is written Na₂SO₄, and not SO₄Na₂, or O₄SNa₂. The atoms occupy a definite position in their arrangement in the molecule. We may express this relative and constant position of the atoms by means of

GRAPHIC FORMULAS.

When we write H₂SO₄ we employ a *rational formula*—i.e., rational, because it shows us in a rational way at a glance how many hydrogen atoms are replaceable and how many bonds the radical has, etc. If we should write O₄SH₂, that would imply that sulphuric acid contains 4 atoms of oxygen, 1 of sulphur and 2 of hydrogen, but it would tell us nothing further than that.

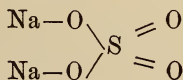
The following graphic formulas of the principal acids should be remembered by the student:



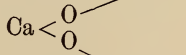
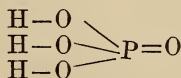
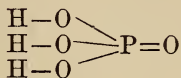
This shows the number of bonds of each element to be satisfied or united with the bond of some other element, and the positions of the atoms in the molecules. In H_2SO_4 , for example, the S has six bonds distributed among the three oxygen atoms, each of which atoms has two bonds. The hydrogen *always* occupies the left-hand position, and in the case of oxyacids is *always* united to oxygen.

Graphic formulas of salts are written exactly as those of the acids are, except that the metals occupy the place of the hydrogen.

Na_2SO_4 would be written graphically thus:



$\text{Ca}_3(\text{PO}_4)_2$ from two molecules of phosphoric acid, H_3PO_4 :



Calcium being a diad and phosphoric radical a triad, we must take three of the former and two of the latter to satisfy all bonds.

It was at one time supposed that all acids contain oxygen, and the element known as oxygen was so termed because of that belief, the term meaning *acid generator*. We are acquainted at present, however, with many bodies which exhibit all the properties of acids and yet contain no oxygen. The combinations of the halogen elements with hydrogen are such bodies—hydrochloric acid, HCl ; hydrobromic, HBr ; hydriodic, HI , etc. These are the *hydracids*. The proportions of the two elements in each of these acids cannot vary as the elements do in the oxyacids. As illustrations of varying quantities of elements in some of the latter, the following may serve:

H_2SO_3 , hydrogen sulphite or sulphurous acid.

H_2SO_4 , hydrogen sulphate or sulphuric acid.

HNO_2 , hydrogen nitrite or nitrous acid.

HNO_3 , hydrogen nitrate or nitric acid.

The oxygen varies in quantity and determines the names of the acids and the salts obtained from them. The student is advised here to read over again what was said on chemical nomenclature and writing of salts elsewhere.

The inorganic acids are never made by the pharmacist; their manufacture on the large scale is an important and one of the most extensive branches of the chemical industry. The manufacturing chemists send the acids into the markets usually in three qualities, the common or *commercial*, used for all else than for medicinal or analytical purposes; the *medicinally pure*, abbreviated *M.P.*, used for internal administration, and the *chemically pure*, abbreviated *C.P.*, intended for analytical purposes where absolute purity is essential. The latter is almost wholly

used in pharmacy for prescription use, and in making preparations containing acids, since it is nearly as cheap as the medicinally pure. One of the most difficult and expensive parts of the manufacture of the acids is the removal of impurities, and for purposes in which the impurities are of no consequence the process of purification is wholly omitted, yielding the common or commercial acids. Very often the presence of minute traces of some kinds of impurities will not affect the value of the acids for medicinal purposes, so that the traces are not removed, but only the greater bulk of the impurities, yielding acids which answer the purposes of medicine very well, but which are yet unfit for delicate analytical uses. For the latter purpose all traces of impurities must be removed, and to do that requires more skill and expense than to free the common acids from the greater bulk of their impurities. In most cases it is not possible to remove all impurities at once from the crude acid, hence the amount of work is greater for the production of the medicinally pure than for the common. The same is true of the chemically pure over the medicinally pure. At one time the cost of the production of the chemically pure was much greater than that of the medicinally pure, but of late, improvements in the methods of purification have brought the price of the former so near that of the latter that pharmacists now usually purchase the *C.P.*

The labels should not satisfy the conscientious pharmacist as to the purity of the acid which he purchases. It is an easy matter to apply the simple tests given in the Pharmacopœia for the detection of impurities. The mineral acids should never be kept in any other than glass-stoppered bottles; cork is acted upon and quickly contami-

nates the acid. A special bottle, called acid-bottle, usually made of green glass and very strong, is used as acid container. The neck and stopper are usually well ground and fit very accurately, for which reason it is often difficult to extract the glass stopper. To do this successfully, all that is usually necessary is to remove the wax or lute from around the stopper and strike the side of the stopper gently with the wooden handle of a spatula, at the same time pulling upward on the stopper. If this will not accomplish the purpose, the stopper may be wedged in between two closet-doors closing against each other, pressing gently inward and cautiously turning the bottle. Great care should be exercised not to break the neck of the bottle, else damage to clothing and fixtures will result from the spilling of the acid. Should this means not avail, the neck of the bottle may be gently turned in the flame of a Bunsen burner for a moment. The heat expands the neck first, and the stopper may be removed before the heat is communicated to it.

The utmost care must be observed in handling the strong acids; some of them—for instance, nitric—may produce fire if poured or spilled upon organic bodies. They are all very corrosive and caustic, attacking and destroying everything they come in contact with. It is always best to have an alkali on hand, in case it becomes necessary to neutralize them; ammonia-water neutralizes most quickly, bicarbonate of sodium or chalk will answer well also. The solutions of soda and potassa neutralize well, but they, too, are very caustic and may do damage if used in excess.

Antidotes to Mineral Acids.—The acids are all very powerful poisons and produce the most painful and dan-

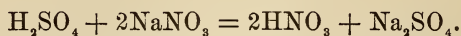
gerous effects on the mucous membrane and stomach. If they are taken at any time accidentally, large draughts of alkaline liquids that do not corrode should be given, followed with oil. Sodium bicarbonate made into a milk with water should be given slowly, if at hand, or any other chemical that will neutralize the acid without itself being injurious. The CO_2 , which is disengaged usually escapes through the œsophagus and mouth, and because of the disengagement of the gas where carbonates are given the latter should not be administered too copiously, but rather in small quantities frequently. The gas may cause undue distention and distress if liberated too suddenly. Soap is an excellent antidote, given in strong solution. The oil, which should follow, obtunds the action of the acid upon the delicate membranes.

Medical Properties of the Inorganic Acids.—Externally applied the inorganic acids are very corrosive, and should be used with great care. For internal administration they are usually given in a diluted form for their refrigerant and tonic properties. They are apt to injure the dental structure by attacking and dissolving the substance of the teeth, and for this reason should always be taken in plenty of water and through a glass tube in such a way that they do not come in contact with the teeth.

Sulphuric Acid.

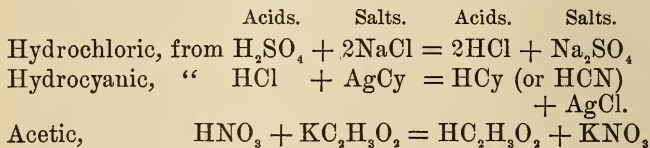
WE know by this time that an acid, aside from its characteristic physical properties, is a chemical compound containing hydrogen as its base, which hydrogen is replaceable by metals. In the same way that we can replace this hydrogen in the acids by metals and produce salts, we can,

reversely, replace the metals in salts by hydrogen and obtain the acids. This will become clearer, perhaps, by illustration. If we treat the salt nitrate of sodium, NaNO_3 , with sulphuric acid, H_2SO_4 , two things happen—the H_2 is replaced by the Na , two atoms of which are taken, and the two atoms of sodium are replaced by an equivalent of hydrogen, H_2 . In other words, there will be simply an interchange of bases; the H_2SO_4 forms the salt Na_2SO_4 , sulphate of sodium, and the nitrate of sodium becomes HNO_3 , nitric acid :



In this manner is nitric acid made. In fact, all of the acids may be made in this way, by bringing an acid in contact with a salt containing the corresponding acid radical.

The following acids, made by acting on the corresponding salts with some other acid, may be cited as examples :



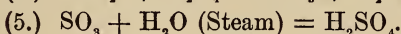
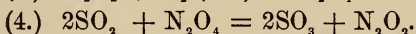
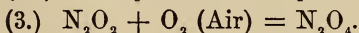
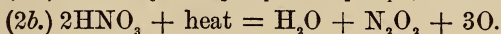
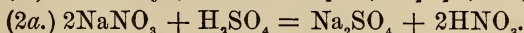
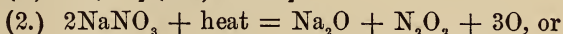
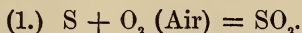
In the above reactions the Na_2SO_4 , AgCl and KNO_3 are *by-products* and the acid the *primary product*.

The question may be put by the student: Are all acids made from some other acid? Must there not be some acid or acids (the source of all other acids) which is or are not made from some other acid? The logical answer would be that if all the acids are made from other acids, very soon there would be no acids at all, and that therefore some acid must have its source in something which is not an acid.

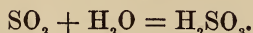
The latter is true of sulphuric acid, which is the most important of all the inorganic acids, principally because it is the source of many other acids, and is not itself made from another acid. The student will better understand this as we go on.

The method of manufacturing sulphuric acid to-day is essentially, with some modifications, the same as that experimentally employed by Lavoisier in 1774 and 1775.

Preparation.—By burning sulphur in air it takes up oxygen and becomes the oxide of sulphur, SO_2 . The heat produced is utilized in suitable chambers to decompose sodium nitrate; this decomposition yields oxygen, which is taken up by the SO_2 to form SO_3 . The SO_3 fumes, coming in contact with watery vapors, unite with them, producing sulphuric acid, H_2SO_4 , in a diluted form. The following equations will serve to demonstrate this more clearly.



If the SO_2 is brought in contact with steam H_2SO_3 , sulphurous acid, will result:



To get the higher form of oxidation, the nitrate of so-

dium is decomposed, the products of decomposition being, besides oxygen and water, an oxide of nitrogen, N_2O_2 , which has the peculiar property of abstracting oxygen from the air to form N_2O_4 and of giving this up again to the SO_2 to make SO_3 , becoming itself again N_2O_2 , when it is again ready to abstract a fresh supply of oxygen from the air. It thus acts as a carrier of oxygen from the air to the SO_2 , and is generated only in the beginning of the process.

Instead of using sulphur, sulphide of iron, FeS , is often employed. In this process it is not necessary that any other acid be employed; it furnishes *the* acid from which nearly all others are made—it is the starting point in the acid-manufacturing industry.

The fumes of SO_2 are passed into large leaden chambers, where they come in contact with the N_2O_4 , and then with steam. The acid thus produced collects at the bottom of the chamber, and is not very concentrated. It forms one of the varieties of acid found in the market, and known as *chamber acid*, of a specific gravity ranging from 1.40 to 1.50. When removed to pans, also of lead, and boiled down to a specific gravity of about 1.70, it constitutes *pan acid*. At that density it begins to act on the lead, the dilute acid having no effect on that metal. In order to further concentrate it, it is necessary to distil off the water in porcelain or platinum or glass retorts.

The sulphate of sodium, which is the by-product, is recrystallized and sold as Glauber's salt. Sulphuric acid thus prepared and purified is an oily liquid of a specific gravity of 1.835, colorless, without odor, of a strongly acid reaction, and is very caustic and corrosive, soluble in alcohol and water with the evolution of heat. If heated to a high

degree it volatilizes; in contact with organic matter it becomes black, carbonizing the body by uniting with the other elements and setting the carbon free.

In the crude acid there are present a number of impurities, which through incomplete methods of purification are often found in smaller quantity in the medicinal and purified acids. To guard against the presence of impurities the pharmacist should always examine the sulphuric acid for lead, copper, iron and arsenic, nitric, hydrochloric and sulphurous acids. In this country the acid is usually made from sulphur, but that made abroad is from sulphides of iron, or iron pyrites, which are always associated with arsenic. The latter, because it is volatile, is usually present in the crude acid made from the iron pyrites, and is a dangerous contamination.

Tests for Impurities in Sulphuric Acid.—The following tests should be memorized by the student, because they are for substances which are not present in sulphuric acid only, but which may be found in the acids prepared directly or indirectly from sulphuric acid.

To detect lead, a portion of the acid should be poured into four volumes of alcohol—a precipitate within an hour indicates the presence of lead.

For iron, when diluted with 10 volumes of water, the addition of excess of water of ammonia will give a brownish precipitate.

Copper, if present, gives a blue color with excess of water of ammonia.

Arsenic may be detected by mixing 1 cc. of a mixture of 1 volume H_2SO_4 and 2 of water, with 1 cc. test-solution stannous chloride, and adding a small piece of pure tin-foil

—if a coloration appears within one hour arsenic is present.

The acid upon supersaturation with ammonia should leave no fixed residue upon evaporation and ignition; this would indicate the absence of non-volatile impurities.

Nitric acid, if present, will produce a brownish or reddish zone between a layer of sulphuric acid, in a test-tube, and one of a freshly-prepared solution of ferrous sulphate.

Hydrochloric acid, if present, gives a precipitate with solution of sulphate of silver.

Uses and Properties of Sulphuric Acid.—The employment of this acid is more extensive than that of any other acid. It is the strongest and most powerful of the official acids, decomposing salts with a marked chemical energy. It is the source of most acids and sulphates, and, in a diluted form, is a valuable medicament. It is sometimes employed externally as an escharotic, but its use for this purpose is limited, owing to its tendency to spread on the skin and to affect the surrounding tissue. When accidentally dropped on the skin its caustic effect may be neutralized by profuse use of diluted water of ammonia, or sodium bicarbonate, or calcined magnesia.

In manipulating sulphuric acid it should be particularly remembered that it evolves a great amount of heat when mixed with other liquids, and that it should *always be poured into the other liquid*, and very slowly, to prevent the vessel from breaking by the sudden expansion caused by the rapid rise of temperature.

The sulphuric acid of the U. S. Pharmacopœia is a liquid composed of 92.5 per cent of absolute H_2SO_4 by weight and 7.5 per cent of water. It should be kept in glass-stoppered bottles. It is colorless, of oily or syrupy consist-

ence, miscible with water and alcohol. It has a specific gravity of 1.835 at 15° C., and is vaporized when heated on platinum-foil

Diluted Sulphuric Acid

contains 10 per cent of absolute acid, and is made by pouring the acid into water, slowly and carefully, stirring meanwhile. It has the same properties as the stronger acid, except those which depend upon the strength or concentration of the latter. Its use is chiefly for internal administration or for dissolving quinine when quinine is desired in the liquid form.

Aromatic Sulphuric Acid.

This is a more grateful preparation for the internal use of sulphuric acid. It contains 20 per cent of absolute acid in *alcohol*, combined with the aromatics ginger and cinnamon. Tincture of ginger and oil of cinnamon are employed by the U. S. P. of 1890. The 1870 Pharmacopœia employed the crude drugs, instead of these their preparations, and the finished product contained, as a result, a little tannic acid. The aromatic, as the dilute acid, is often employed for preparing solution of quinine; but the presence of tannic acid in the 1870 Pharmacopœia product frequently caused a turbidity, from the fact that tannic acid precipitates alkaloids, quinine in this case. To obviate this, the Committee of Revision of the Pharmacopœia in 1880 substituted tincture of ginger for the powdered root, and oil of cinnamon for the powdered bark, as mentioned above.

FUMING SULPHURIC ACID.—The sulphuric acid described above is the ordinary acid of commerce. There is another variety known as *Fuming Sulphuric Acid*, or *Nordhausen acid*, so called because it constantly emits fumes when ex-

posed to the air; and because it was first made at a place called Nordhausen, in Saxony, Germany. This acid is made by distilling, in large stone retorts heated to redness, the ordinary sulphate of iron previously deprived of its water of crystallization by heat. Polishing-rouge remains in the retort, and the fuming acid distils over. It is supposed to be $\text{H}_2\text{S}_2\text{O}_7$ or $\text{H}_2\text{SO}_4\text{SO}_3$ —that is, a solution of SO_3 fumes in H_2SO_4 . The SO_3 fumes are constantly given off.

Solid Sulphuric Acid

is the anhydride, SO_3 , now to be found in a solid condition in the market is soldered boxes of tinned sheet-iron. When moisture is wholly excluded this acid has little or no action on metals.

Sulphurous Acid.

THE termination of the first word in this title indicates that the acid is a lower sulphur acid than sulphuric acid; it denotes that it is the next lower. If we were to take one atom of oxygen away from sulphuric acid, H_2SO_4 , the acid H_2SO_3 would result, and in practice that is the precise way in which sulphurous acid is made.

Preparation.—Sulphuric acid is mixed with powdered charcoal and heated in a retort or large flask, and the generated gases passed through two bottles containing water, the last of which will contain the sulphurous acid sought. The first bottle washes the gas. The charcoal or carbon acts on the sulphuric acid and takes away one atom of oxygen from each of two molecules of acid. The H_2SO_3 which is left splits up into H_2O and SO_2 ; the SO_2 and CO_2 pass through the water in the bottles, saturating the water with

SO₂, the CO₂ escaping. A third bottle containing a solution of carbonate of sodium is sometimes employed to retain any gas that may not be absorbed by the distilled water in the other bottles. The solution of SO₂ in the second bottle constitutes the product which the process was intended to furnish, namely H₂SO₃. The first bottle contains the same but mixed with other bodies, and is rejected. The steps may be illustrate by the following equations:

1. $2\text{H}_2\text{SO}_4 + \text{C} = \text{CO}_2 + 2\text{H}_2\text{SO}_3$.
2. $\text{H}_2\text{SO}_3 + \text{heat} = \text{H}_2\text{O} + \text{SO}_2$.
3. $\text{SO}_2, \text{ cold, } + \text{H}_2\text{O} = \text{H}_2\text{SO}_3$.

The strength of the acid is 6.4 per cent sulphurous acid gas, SO₂. The solution of sodium carbonate absorbs the excess of SO₂ which the water will not absorb: $\text{Na}_2\text{CO}_3 + \text{SO}_2 = \text{Na}_2\text{SO}_3 + \text{CO}_2$, liberating more carbonic acid gas.

Properties.—This solution of sulphur dioxide is a colorless liquid of a sp. gr. not less than 1.035, strongly acid reaction, and an odor resembling burning sulphur. When heated it completely volatilizes. Its bleaching properties are pronounced, and depend upon the SO₂ which it contains.

It is rarely used internally, its salts being preferred. Dose, from 3 to 30 minims (0.185 to 1.85 cc.) or more, with water.

Its chief property is anti-parasitic, destroying microscopic organisms, for which reason it may be employed when an anti-ferment is desired. Its employment to prevent fermentation is manifold, that to check fermentation of cider at a certain stage being very extensive.

Nitric Acid.

THE acid next in importance to sulphuric acid is probably nitric acid, HNO_3 , a combination of the highest oxide of nitrogen, N_2O_5 , with water. It may be looked upon as consisting of $2\text{HNO}_3 = \text{H}_2\text{O} + \text{N}_2\text{O}_5$. Nitrous acid contains in two molecules the same amount of water united with the next lower, or nitrous, oxide, $\text{N}_2\text{O}_3 - 2\text{HNO}_2 = \text{H}_2\text{O} + \text{N}_2\text{O}_3$.

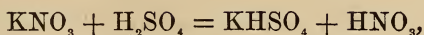
Preparation.—Nitric acid is made on a large scale by acting with sulphuric acid on either sodium or potassium nitrate contained in a retort, heating carefully to distillation and collecting the distillate in well-cooled receivers. Equal weights of the acid, and nitrate in case of the potassium salt, are usually employed, but in some processes on a large scale more of the nitrate and less of the acid are preferably employed.

These two reactions may be illustrated by equations, and are a means of illustrating

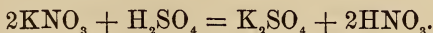
THE FUNCTIONS AND USES OF EQUATIONS.

An equation tells not only what will be produced when two bodies capable of uniting chemically are brought in contact, but it tells also how much by weight of each body must be employed, or, on the other hand, how much by weight of each of the resulting compounds will be formed, or how much of each of the reacting bodies must be used to produce a certain weight of one or more of the resulting compounds.

The first of these reactions is brought about when we take weights represented by one molecule of each, acid and salt, as follows:



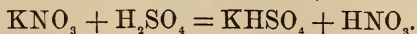
and the second, when twice as much of the nitrate is employed:



In the first reaction one molecule of the nitrate is used and one molecule of nitric acid is produced, besides an *acid* sulphate of potassium, in which there is still some unreplaced hydrogen. In the second case two molecules of nitrate of potassium are used and two of nitric acid are produced, with the formation of a *neutral* sulphate of potassium as a by-product. The first process is employed when an exceedingly concentrated acid is not demanded; the acid sulphate, too, is more easily removed from the retort than the neutral sulphate, which often forms a very hard cake.

Now, to return to equations and their functions, we find that a *symbol* represents, among other things, the weight of the atom, or *atomic weight*; thus, K means, besides other things, that the potassium atom is 39 times heavier than the hydrogen atom. The *formula of a compound* represents, besides other things, the weight of the molecule—that is, the sum of its atomic weights; thus, KNO_3 means that the weight of the molecule is 39 for K, + 14 for N, + 48 for O_3 = 101. In the same way we find that the molecular weight of sulphuric acid is 98. Now, these figures denote the proportion in which the bodies whose molecular weights they represent will unite with each other, according to the first equation:

$$101 + 98 = 136 + 63.$$

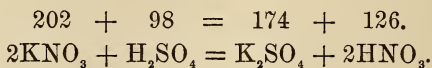


That means if 101 of nitrate of potassium and 98 of acid

be used, 136 of acid sulphate and 63 of nitric acid will be produced.

The figures denote *relative proportions by weight*. They may mean grammes, ounces or pounds, or any other denomination of weight, but they must all represent *one kind* of weight division.

According to the second equation, different proportions of nitrate and acid would enter the reaction, as follows:



In the chapter on chemical philosophy we learned that elements united with each other in definite chemical proportions, and that sometimes the proportion of one element is fixed, while the other increases in definite *multiple proportions*. Thus in CO and CO₂ the C is fixed or constant, while the O increased in the case of CO₂ in multiple proportion of 16—namely, twice 16, etc. This is not true of elements alone, but of compounds as well. In the above two reactions the acid is constant, but the KNO₃ may be used in more than one proportion, subject to the law that it must be exactly a multiple proportion, and no more nor less than such multiple proportion expresses, the proportion being always understood to be taken by weight.

So far we have learned that if we express the reactions between two or more bodies by equations, we can find out by employing the molecular weights of the bodies entering the change how much of each of those bodies must be employed without causing a loss of either.

We may now proceed a step further in our chemical mathematics. Writing an equation as we did above, we can tell from it how much of the bodies entering the reac-

tion must be used to produce a *desired amount of the product*.

Thus, supposing it was our intention to make 100 pounds of nitric acid. From the above equation we know that in order to produce 63 pounds we must employ 101 pounds of the potassium salt and 98 pounds of sulphuric acid, providing both were absolute—that is, 100 per cent. The proportions necessary for a given quantity may be calculated by the rule of three thus: If 63 pounds metric acid require 101 pounds of nitrate, then 100 pounds would require x quantity; or, as 63 is to 100, so must 101 be to the unknown quantity:

$$63 : 100 :: 101 : x.$$

$$63)10100(160.31+$$

= number of pounds of the nitrate necessary to produce 100 pounds nitric acid.

$$63 : 100 :: 98 : x.$$

$$63)9800(155.55+$$

= number of pounds of sulphuric acid necessary to make 100 pounds of nitric acid.

When we say that, according to the equation, 98 pounds of sulphuric acid must be employed, we mean 100 per cent acid. If the acid were only 96 per cent absolute acid, we would need to employ of it more than 98 pounds. The weight of the 96 per cent acid to be employed would be ascertained in this wise, simply:

$$96 \text{ per cent} : 100 \text{ per cent} :: 98 \text{ pounds} : 102.2 \text{ pounds}.$$

Answer.—102.2 pounds of the 96 per cent. sulphuric acid, therefore, contains the necessary 98 pounds of absolute acid.

Properties of Nitric Acid.—This acid is, next to nitro

hydrochloric, the most caustic of all acids. It is extremely corrosive, and has a powerful oxidizing capacity, greater than any other body. It was made as early as the thirteenth century, and was called *aqua fortis* by the ancients. It is now called nitric acid because made from *nitre*, the common name for potassium nitrate. Nitric acid is known also by the name *azotic acid*.

By heat it is wholly volatilized; a residue would indicate presence of fixed impurities. It usually has a straw-yellow color, due to the presence of a lower acid of nitrogen, but when perfectly pure it is wholly colorless. Its action upon living tissue is that of a strong caustic; it stains the skin an indelible yellow color, producing what some call xanthoproteic acid. The most characteristic property of nitric acid is its oxidizing powers; it oxidizes phosphorus, sulphur, etc., and dissolves most of the metals; with salifiable bases it forms nitrates. It enters into the preparation of nitrohydrochloric acid, the only acid that will dissolve gold. In its action on many organic bodies its chemical affinity exerts itself with such violence as to produce very often enough heat to ignite the surrounding bodies. Its oxidizing properties depend on the evolution of nascent oxygen in its decomposition: 2HNO_3 decomposes into $\text{H}_2\text{O} + \text{N}_2\text{O}_2 + 3\text{O}$. The specific gravity of the official acid is 1.414, corresponding to a strength of 68 per cent absolute acid; but the acid usually furnished by the manufacturing chemist is seldom of that strength, being more often of a specific gravity of 1.35. Stress should always be laid on the importance of employing the 1.414 specific gravity acid in making the U. S. P. preparations, as that was used in the calculations for the preparations into which nitric acid enters.

Impurities in Nitric Acid.—All the impurities that may be found in sulphuric acid may be looked for with the appropriate tests given under sulphuric acid. In addition, bromic and iodic acids may be present, which may have been in the nitrate as iodate and bromate. If the acid be diluted and shaken with a few drops of chloroform, the latter should not become colored—this would indicate the absence of iodine or bromine. The introduction of a small piece of pure zinc produces no color in the absence of iodic or bromic acids.

Uses in Pharmacy and Medicine.—Nitric acid largely diluted is the official dilute acid, which is given in doses of 5 to 15 minims (0.30 to 0.90 cc.) as a tonic and astringent. In its undiluted form it is used as an escharotic for nævus, warts, etc.

Dilute Nitric Acid

Has the same properties and uses as the strong, excepting those depending on greater concentration. It contains 10 per cent of absolute acid.

Hydrochloric Acid.

IN pharmaceutical nomenclature there is a certain specific meaning implied by the titles occurring in the Pharmacopœia, distinct altogether from the ordinary or usual meaning attached to those words. Thus the title "*Acidum Hydrochloricum*," or its English equivalent, "Hydrochloric Acid," means the particular body described in the U. S. P. under this head, and no other. There are other *kinds* of hydrochloric acids, differing not in the essential of containing the elements hydrogen and chlorine, but in strength, in the presence of impurities and contaminations, etc.

The hydrochloric acid, therefore, that we are to consider is "a liquid composed of 31.9 per cent of absolute acid and 68.1 per cent of water," as described in the Pharmacopœia. It is at times called "muriatic acid," "spirit of salt," "marine spirit," "marine acid," the latter terms referring to its source, common salt from the sea.

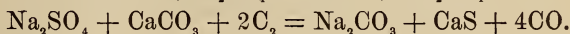
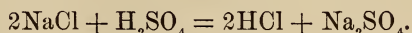
Properties.—The pure acid is colorless; has a penetrating, suffocating odor and extremely caustic taste, miscible with water and alcohol in all proportions. It has a specific gravity of 1.16, but may be concentrated to increase its density up to 1.22. Heat will completely volatilize it, and even at ordinary temperatures it slowly vaporizes, forming white fumes of ammonium chloride by combining with the ammonia in the air. This is a reliable test for either hydrochloric acid or ammonia; if a glass rod be dipped into ammonia water and held near the acid, the white fumes immediately appear. Ammonia, $\text{NH}_3 + \text{HCl} =$ ammonium chloride, NH_4Cl . When very concentrated the acid blackens organic matter, as does sulphuric acid. The acid is not a liquid, but a gas, HCl , the only known combination of hydrogen and chlorine. When dissolved in water to the extent of 31.9 per cent, it constitutes the official acid.

The acid and any of its salts form with solution of silver nitrate a characteristic white precipitate of chloride of silver, which is readily soluble in ammonia, but reprecipitated if the ammonia is again neutralized with an acid. Hydrochloric acid has strong chemical properties, combines with the alkalies and alkaloids to form salts, chlorides, and dissolves a number of the metals.

Preparation.—If common salt, chloride of sodium, be mixed with sulphuric acid and the mixture distilled, hydro-

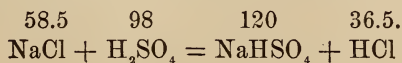
chloric acid or the chloride of hydrogen will result. Unless the salt and acid are pure the product will be contaminated. On a large scale a very impure acid is obtained by the employment of common sulphuric acid and distilling from iron retorts.

By far the greater part of the hydrochloric acid of the market is not made directly, but is a by-product in the manufacture of sodium sulphate in the process of converting the chloride of sodium into carbonate and hydroxide, or into "soda-ash," as the mixture is called. Chloride of sodium is the source of most of the other sodium salts, which latter, however, cannot be made directly from the chloride. The chloride is first changed into the carbonate, from which most of the other sodium salts are then made with facility. To convert the chloride into the carbonate, Leblanc's process is one means, and in the first stage of the process hydrochloric acid is a by-product:

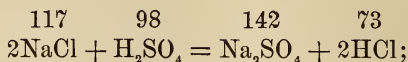


The salt NaCl reacts with the sulphuric acid and produces hydrochloric acid, which is in a gaseous condition, and is made to pass through towers, down which water trickles to dissolve the gas, and the acid then passes out as a relatively concentrated product.

To produce a pure acid, pure fused chloride of sodium and pure sulphuric acid are employed, and the gas passed into distilled water contained in a series of bottles. The following reactions illustrate the chemistry of the process. Here, as in nitric acid, two proportions may be employed:

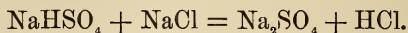


or,



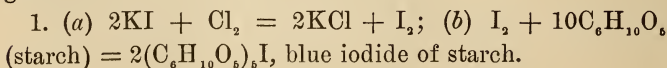
The figures denote the molecular proportions of NaCl and H_2SO_4 to be employed, as fully explained under sulphuric acid.

The first process yields acid sulphate of sodium, which is more easily removed from the retort than the neutral sulphate obtained in the first reaction. The acid sulphate may be further employed with more NaCl to yield HCl:



Converting the sulphuric acid first into the acid sulphate, and that into the normal or neutral, amounts to the same thing as does the second reaction above.

Impurities.—As in nitric acid, so in hydrochloric acid may be present such impurities as are in the sulphuric acid from which it is made. The tests given under sulphuric acid may be here applied for these impurities. Additional impurities or contaminations may be chlorine and sulphuric acid. Chlorine will be recognized by liberating the iodine in a solution of iodide of potassium, the acid having previously been diluted with five or six volumes of water. The iodine will be recognized by the familiar starch-test—that is, turning blue with starch-paste. If chloride of barium solution produces a precipitate when added to the dilute acid, the presence of sulphuric acid would be indicated. The following reactions illustrate the two tests just given:



2. $\text{BaCl}_2 + \text{H}_2\text{SO}_4 = 2\text{HCl} + \text{BaSO}_4$, heavy white insoluble sulphate of barium.

(It is doubted by many that iodide of starch is a definite chemical compound. The Pharmacopœia of 1880 evidently did not recognize it as such, to judge from the name it gave to it—*iodized* starch.)

The commercial acid is full of many kinds of impurities, the most injurious being sulphurous acid, which is present sometimes to the extent of 10 per cent. The pharmacist should never sell it except for use in the arts.

Uses of Hydrochloric Acid.—This acid is used primarily for the preparation of the dilute acid and the nitrohydrochloric acid, but also largely as a medicine for its tonic, refrigerant and antiseptic properties. The dose is about five minims, largely diluted. In concentrated form it is a caustic, though not so powerful as nitric acid. It is frequently employed to dissolve alkaloids, iron and zinc in making the respective solutions of the Pharmacopœia, and also aids in dissolving the arsenic in the official solution of arsenous acid. With black oxide of manganese it yields the official chlorine-water, and by dissolving out the phosphate of calcium from animal charcoal it produces the purified animal charcoal. With cyanide of silver it gives diluted hydrocyanic acid.

Antidote.—Hydrochloric acid is an irritant poison, producing a characteristic fiery redness of the tongue, blackness of the lips and violent gastric pains. The best antidote is calcined magnesia made into a milk with water, given in copious draughts. Other mildly alkaline solutions, as soap, etc., are also efficient. Demulcent and bland drinks should follow.

Dilute Hydrochloric Acid.

This acid has the same properties as the strong acid, from which it is made, excepting those depending on greater concentration. It contains 10 per cent of absolute acid, and has a specific gravity of 1.050.

Nitrohydrochloric Acid.

THIS acid, sometimes called "nitromuriatic acid," "chloronitric acid," "chloroazotic acid," "*aqua regia*," etc., is made by mixing 18 parts of nitric acid with 82 parts of hydrochloric acid, in a capacious vessel, and allowing to stand until effervescence ceases. It is then directed to be poured into glass-stoppered bottles, which should not be more than half filled, and kept in a cool place. This acid was called *aqua regia* by the early chemists because of its property of dissolving gold, the king of the metals (*regia* from *rex*, which means king). There are various opinions regarding the nature of the reaction between strong nitric and hydrochloric acids, but the nitrosyl chloride theory seems to be shared by most authorities at present: $\text{HNO}_3 + 3\text{HCl} = 2\text{H}_2\text{O} + \text{NOCl} + \text{Cl}_2$. The presence of free chlorine gives, very probably, to the acid its power to dissolve gold and other metals which have little affinity for oxygen. The nitrosyl chloride, NOCl , remains entirely inactive and unchanged during the solution of gold.

Nitrohydrochloric acid is a golden-yellow, fuming, corrosive liquid, having a strong odor of chlorine; it liberates iodine from solutions of iodides. If kept exposed to sunlight it decomposes into hydrochloric acid, principally, and loses its chlorine. The pharmacist should never keep it in large quantity because of its proneness to decomposition

and deterioration. Especial care should be observed not to transfer it to bottles before effervescence has entirely ceased, lest the pressure burst the bottles.

Uses in Medicine.—Nitro-hydrochloric acid is employed in diseases of the liver, stimulating especially the flow of bile. It is used externally in very dilute solution as bath. The dose of the fresh acid is 3 to 5 drops, largely diluted.

Dilute Nitro-hydrochloric Acid.

Made as the above, with the addition of water *when effervescence has entirely ceased*; the dilute nitric and hydrochloric acids do not react, hence the employment of the strong acids and subsequent dilution with water. It is claimed by good authority that the water in the dilute acid causes a decomposition of the products formed by the reaction of nitric and hydrochloric, and a re-formation of the original acids, with a little nitrous acid as well.

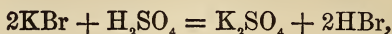
Clinical experience confirms the claim that dilute nitro-hydrochloric acid is not an eligible preparation. The strong acid is employed preferably, well diluted just before administration.

Dilute Hydrobromic Acid.

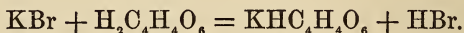
Properties.—Hydrobromic acid is a colorless, clear liquid, of a specific gravity of about 1.077, wholly volatile upon application of heat, odorless and of strongly acid taste and reaction. It contains 10 per cent by weight of absolute acid, corresponding in strength to most of the other dilute acids. There is no stronger acid official, but the market affords several that are more concentrated.

Preparation.—The Pharmacopœia does not give a process

for the preparation of this acid, but there are numerous methods in vogue for its production. The following yields a pure product: Sulphuric acid is made to react on an equal quantity of bromide of potassium in solution, and allowed to stand, to permit the sulphate of potassium, which is formed, to crystallize. The liquid is then poured off into a retort and distilled. The *rationale* of the process is this:



which, it will be seen, is the identical process employed in making hydrochloric acid, excepting that, in making the latter, distillation is resorted to at once. Another method, known as the "precipitation process," consists in reacting with tartaric acid upon potassium bromide, and allowing the acid tartrate of potassium, or cream of tartar, to crystallize:

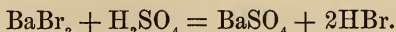


If the mixture is put away for several weeks in a place removed from dust and jarring, nearly all the cream of tartar will crystallize out, leaving the HBr in the liquid.

The addition of a little alcohol facilitates the separation of the acid tartrate. This is known as Wade's Buchanan's method. *Fothergill's acid* is made after Wade's method, but it is somewhat weaker in strength. The acids made by these methods are open to the objection of containing small quantities of cream of tartar in solution.

Among the many other methods only one more need be mentioned here, that of Goebel's Glover's process, which consists in treating barium bromide, BaBr_2 , with sulphuric

acid, producing HBr and the insoluble barium sulphate as precipitate:



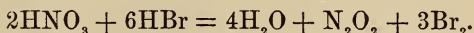
The first of these methods, known as Squibb's process, has the advantage over the others of producing a product of greater purity and more definite strength.

Tests of Identity.—Solution of silver nitrate produces a yellowish-white precipitate, which is insoluble in nitric acid and very sparingly soluble in ammonia.

If chlorine or nitric acid be added to the acid, bromine will be liberated, which is soluble in chloroform or disulphide of carbon, giving to these liquids a yellow color.

The chlorine has a greater affinity than the bromine has for the hydrogen in the HBr, and it therefore removes the bromine to take its place: $2\text{HBr} + \text{Cl}_2 = 2\text{HCl} + \text{Br}_2$.

Nitric acid, when it liberates the bromine, probably does so in this wise:



Impurities.—Upon keeping for any length of time the acid sometimes becomes yellow; this is due to the liberation of bromine. It should not be used unless wholly colorless.

Sulphuric acid may be present; if so, it may be detected by BaCl_2 , which gives an insoluble white precipitate.

Uses.—Dilute hydrobromic acid is employed medicinally for the same purposes that bromide of potassium and bromide of sodium are used. It is claimed to have advantages over the alkali bromides, as potassium and sodium are said to exert a depressing influence upon the system. The dose is up to 2 fluid drams (7.39 cc.) given in syrup or water.

Hydriodic Acid.

The third of the halogen acids, hydriodic acid, is official in the form of a *syrup* of 1 per cent strength by weight. This preparation was introduced in the U. S. P. of 1880, and its introduction grew out of the difficulty of keeping hydriodic acid in aqueous solution, the iodine always becoming liberated, in a short time rendering the preparation unfit for administration. Experiments have shown that syrup will efficiently preserve the acid. There is about 1.30 gramme acid in 100 cc. syrup.

Preparation.—Several methods may be employed in the preparation of *syrup* of hydriodic acid, and each involves a chemical process. One method is as follows: Ten parts of iodine are dissolved with *gentle* heat (extreme heat would produce loss of alcohol by evaporation and of iodine by vaporization) in 80 parts of alcohol. When dissolved the solution is added to 150 parts syrup and 150 parts of water, and a current of hydrosulphuric acid gas is passed through the mixture until it acquires a yellowish color and ceases to assume a brown tinge on shaking. The liquid is then filtered and evaporated during constant stirring at a temperature not exceeding 131° F., until all odor of hydrosulphuric acid gas has disappeared. When cold it is flavored with spirit of orange 5 parts, and 500 parts of sugar added and enough water to make 1000 parts (all “parts by weight”). It is preserved in small, well-stoppered bottles and kept in a cool and dark place. The chemistry of the process is: $I_2 + H_2S = 2HI + S$. The H_2S may be generated from iron sulphide and sulphuric acid: $FeS + H_2SO_4 = FeSO_4 + H_2S$. The FeS is contained in a flask with a double-perforated cork, through which pass a

funnel to the bottom, to convey the acid, and a delivery tube, to carry away the gas. The gas is usually washed by being passed through water before it is conducted into the iodine solution. The sulphur which precipitates during the process remains behind on the filter. H_2S is soluble to a certain extent in water, from which it can be expelled again by heat, hence that direction in the process.

Properties.—Syrup of hydriodic acid is a transparent liquid, sometimes colorless, but usually of a pale-straw color, of a specific gravity of 1.313 and containing 1 per cent of absolute acid.

If the substances used in its preparation are pure, as they should be, it will contain no impurities, but the U. S. P. gives tests for free iodine, chloride and sulphate.

Free iodine will be recognized by the blue color with starch-paste. Chlorides or hydrochloric acid, with silver nitrate give precipitates which are soluble in ammonia. The precipitate with a bromide and silver nitrate is sparingly soluble in ammonia.

Sulphuric acid is detected by $BaCl_2$.

Uses.—This preparation is possessed of the chemical properties peculiar to iodides in general. Its dose is from 15 to 40 minims, considerably diluted.

Phosphoric Acids.

Preparation.—Phosphoric acid, as the official phosphoric acid is termed by chemists, is never made by the pharmacist. By the manufacturing chemist it is made by carefully dissolving phosphorus in dilute nitric acid on a water bath. Dilute nitric acid exerts no action upon phosphorus in the cold, but at an elevated temperature it changes it into phosphorous or phosphoric acid, according to its

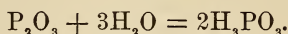
strength, which must be regulated accordingly. The stronger the nitric acid the more rapidly will it oxidize the phosphorus into phosphoric acid, but with the strong acid the reaction is so violent that the acid is always diluted with water for the operation. Serious accidents have happened from explosions resulting from the use of strong acid. The following equations will serve to demonstrate the reactions of nitric acid with phosphorus:

First, when the nitric acid is well diluted :

$2P + 2HNO_3 + 2H_2O = 2H_3PO_3 + N_2O_2$; phosphorous acid is produced. This reaction may be resolved thus:

The HNO_3 , when it oxidizes, as has been mentioned, resolves itself into:

$2HNO_3 = H_2O + N_2O_2 + 3O$; the oxygen attacks the phosphorus, $2P + 3O = P_2O_3$, forming the phosphorous oxide, which with water produces the phosphorous acid:



Second, when the nitric acid is less dilute:

$6P + 10HNO_3 + 4H_2O = 5N_2O_2 + 6H_3PO_4$, forming phosphoric acid. The several steps in this reaction are as in the preceding, excepting that more nitric acid, relatively, is employed, and consequently more oxygen is set free to produce the higher form of oxidation.

$10HNO_3 = 5H_2O + 5N_2O_2 + 15O$; the oxygen again attacks the phosphorus: $6P + 15O = 3P_2O_5$, producing phosphoric oxide, which with water gives phosphoric acid.



The acid last described is official in the U. S. Pharmacopœia.

Varieties of Phosphoric Acids.—Orthophosphoric anhydride, or anhydrous phosphoric acid, P_2O_5 .

Metaphosphoric acid, or glacial phosphoric acid, HPO_3 .

Orthophosphoric acid, H_3PO_4 .

Diphosphoric acid or pyrophosphoric acid, $H_4P_2O_7$.

Besides these there are, among others of less importance, the following:

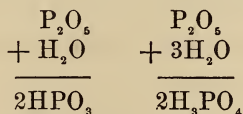
Anhydrous phosphorous acid, P_2O_3 .

Phosphorous acid, H_3PO_3 .

Hypophosphorous acid, H_3PO_2 .

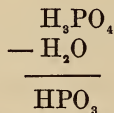
The P_2O_5 is formed when phosphorus is burned in excess of oxygen, and is in form of a white flocculent powder.

With one molecule of water it gives *meta*-phosphoric acid, and with 3 molecules of water *ortho*phosphoric acid.



By adding one molecule of water to one of metaphosphoric acid the orthophosphoric will be formed: $HPO_3 + H_2O = H_3PO_4$. This may be made practically by boiling HPO_3 with water.

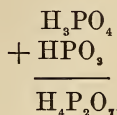
If, on the other hand, orthophosphoric acid be heated to drive off one molecule of water, the metaphosphoric will be produced:



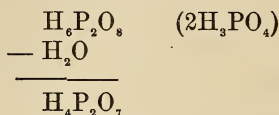
The ortho-acid may be looked upon as a trihydrate, $P_2O_5 + 3H_2O$ ($2H_3PO_4$), while the meta-acid is a monohydrate $P_2O_5 + H_2O$ ($2HPO_3$).

Intermediate between the ortho and meta acids is another, the pyrophosphoric acid, $\text{H}_4\text{P}_2\text{O}_7$.

Theoretically this may be looked upon as consisting of one molecule of orthophosphoric and one of metaphosphoric acids:



or, as being derived from two molecules of orthophosphoric acid by the abstraction of one molecule of water:



The meta- is easily distinguished from the orthoacid by its property of coagulating albumen and precipitating with soluble salts of silver and barium, which properties are not possessed by the orthophosphoric acid. The latter gives a precipitate with silver nitrate in an excess of an alkaline solution only, which precipitate is *yellow*, and different from the precipitate formed by *pyrophosphoric* acid under like conditions, which is *white*. The *pyrophosphoric* acid will not coagulate albumen, wherein it also differs from metaphosphoric acid.

Meta means *beyond*, ortho *straight*, and pyro *fire*. The pyro-acid is so termed because it is usually made by heating the orthophosphoric acid. Heated to a much higher temperature, the pyrophosphoric acid is converted into the meta-acid, and the latter, by boiling in water, is reconverted into the orthophosphoric.

According to the foregoing explanation these three acids

are convertible into each other by the addition or abstraction of water. The salts which these acids produce through the replacement of their hydrogen by metals have the corresponding names:

Metaphosphoric acid, HPO_3 , gives metaphosphates: NaPO_3 , sodium metaphosphate.

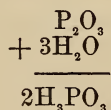
Orthophosphoric acid, H_3PO_4 , gives phosphates: Na_3PO_4 , sodium phosphate.

Pyrophosphoric acid, $\text{H}_4\text{P}_2\text{O}_7$, gives pyrophosphates: $\text{Na}_4\text{P}_2\text{O}_7$, sodium pyrophosphate.

The Pharmacopœia recognizes orthophosphoric acid under phosphoric acid and dilute phosphoric, while the pyrophosphoric acid is represented by pyrophosphate of sodium, $\text{Na}_4\text{P}_2\text{O}_7$, and by pyrophosphate of iron, $\text{Fe}_4(\text{P}_2\text{O}_7)_3$. (The latter is insoluble, but the U. S. P. salt is in form of soluble scales, made so by the addition of citrate of sodium.)

These acids are all derivatives of the phosphoric oxide, P_2O_5 , and have respectively one, three and four replaceable hydrogen atoms.

Varieties of Phosphorous Acids.—The phosphorous oxide, P_2O_3 , yields, as above stated, phosphorous acid, by the addition of water:



This acid forms phosphites with bases. It has only two replaceable hydrogen atoms, and is sometimes written H_2HPO_3 , to show that there are only two replaceable hydrogen atoms. Sodium phosphite would be written Na_2HPO_3 .

The oxide corresponding to *hypophosphorous* acid, H_3PO_2 , is not known ; it would be P_2O .

Hypophosphorous acid has only one replaceable hydrogen atom, and should be written HPH_2O_2 .

It is represented in the Pharmacopœia by sodium, potassium, calcium and iron hypophosphites. Their respective formulæ are NaPH_2O_2 , KPH_2O_2 , $\text{Ca}(\text{PH}_2\text{O}_2)_2$ and $\text{Fe}_3(\text{PH}_2\text{O}_2)_6$.

(The student is advised to become thoroughly acquainted with the chemistry of phosphorus and its acids; as set forth in the foregoing chapter. He will find it necessary to have frequent recourse to it in his practice and studies.)

Properties of Phosphoric Acid.—Phosphoric acid is a clear, colorless, syrupy liquid of strongly acid taste and reaction, containing 85 per cent of absolute acid and 15 per cent water. It is odorless, is miscible with water, and has a specific gravity of 1.710 at 15°C .

Impurities and Tests.—Occasionally small quantities of metaphosphoric acid, phosphates of calcium and magnesium, arsenic, sodium, potassium; hydrochloric, nitric, phosphorous and sulphuric acids, and more than 15 per cent of water are found. Phosphate of sodium has at times been detected in large quantities.

To detect calcium and magnesium, the acid is supersaturated with ammonia-water; a precipitate indicates presence of the bases.

Arsenic, lead and other metals are precipitated by passing sulphuretted hydrogen through the acid for half an hour. A yellow precipitate indicates arsenic.

Sodium or potassium may be detected by precipitating with acetate of lead (producing insoluble phosphate of

lead and soluble acetate of sodium and potassium), filtering and evaporating the solution to dryness. The residue will impart a yellow color to the colorless flame of a Bunsen-burner (sodium), and a purple flame if viewed through blue glass (potassium).

Hydrochloric, nitric and sulphuric acids may be looked for by the tests given before.

Phosphorous acid, if present, decolorizes solution of permanganate of potassium.

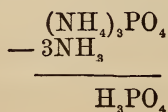
THE PHYSIOLOGICAL ACTION OF PHOSPHORIC ACID is that of a powerful stimulant to the brain and nerves. Authorities differ as to its medicinal value. There is probably some truth in the assertion of J. A. Thompson, that the complete oxidation of phosphorus impairs, or even destroys, its therapeutic properties.

Dilute Phosphoric Acid.

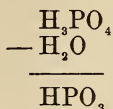
This contains 10 per cent of orthophosphoric acid, and all that has been said concerning phosphoric acid is applicable to the dilute form.

Glacial Phosphoric Acid.

This acid is not made directly from phosphorus, but from calcined bones, by digesting them with sulphuric acid. The insoluble sulphate of calcium which is produced is removed by filtration, the filtrate neutralized with ammonia, separated from the precipitated phosphates and evaporated to dryness. The residue, composed of phosphate of ammonium, is then heated until all ammonia is driven off:



The orthophosphoric acid, H_3PO_4 , upon continued application of heat parts with 1 molecule of water to form the dry fused glacial acid, which is the metaphosphoric acid:



Properties.—Glacial phosphoric acid is in transparent, colorless, glass-like masses, or in the form of sticks, obtained by fusing and pouring into moulds.

It deliquesces on exposure to the air, dissolves in water and alcohol; the solutions having a strongly acid taste and reaction. It is never chemically pure, always containing sodium or other bases, which latter, however, are necessary to the formation of a solid acid. The chemically pure acid is a soft, gum-like mass.

Impurities.—The impurities likely to occur in glacial phosphoric acid are those occurring in phosphoric acid. For the detection of these the same tests may be employed.

Arsenous Acid.

THE substance described in the Pharmacopœia under this title is *not* an acid, all acids, as before stated, requiring hydrogen in the base. The term is a conventional one: it *ought* to be *arsenous oxide*, which better expresses its chemical composition. This *oxide* becomes an *acid* (like SO_2 , CO_2 , P_2O_5 , etc.) when combined with water:



There is another oxide of arsenic, a higher one, As_2O_5 , which with water produces arsenic acid:



Arsenic acid is usually made by oxidizing arsenous oxide with nitric acid:



Preparation.—The arsenous acid is the only one of importance to the pharmacist. It is obtained in nature from arsenical ores by roasting in furnaces and conducting the hot vapors to surfaces on which they may cool and condense, usually cast-iron vessels with conical heads. Arsenic is easily volatilized, taking up oxygen from the air as soon as it enters the gaseous condition, forming As_2O_3 , which condenses into a sublimate.

Properties.—It is a heavy crystalline or amorphous powder, or crystalline masses, semi-transparent, with streaked appearance. Soluble in from 30 to 80 or 90 parts of water, its solubility varying probably with the nature of its structure. In alcohol it is very little soluble, but hydrochloric acid facilitates its solution, as do also the alkalies. Arsenous acid is permanent at ordinary temperatures, but when heated it volatilizes unchanged, emitting a garlicky odor, especially when reduced on ignited charcoal. Chemically, arsenous acid responds to all the tests for arsenic. The latter will be given in the chapter on Arsenic Salts.

Poisonous Properties of Arsenous Acid.—See Arsenic Salts.

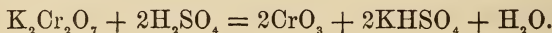
Antidotes to Arsenous Acid.—There are two antidotes recognized by the U. S. Pharmacopœia: (1) the freshly precipitated hydroxide of iron, obtained by adding water of ammonia to solution of tersulphate of iron, pouring precipitate upon muslin strainer, draining and washing with water to remove any excess of ammonia or iron; (2) hydrated oxide of iron with magnesia, prepared by mixing calcined

magnesia with solution of tersulphate of iron. (See chapter on Arsenic Salts for equations.) Until the antidote can be obtained, the poison should be removed as far as possible by inducing vomiting by means of the finger or a feather, or by administering a prompt emetic of mustard, salt and warm water, or 20 grains of powdered ipecac. Demulcent drinks, such as milk, flour and water, or white of egg and water, envelop the poison and retard its solution.

Uses of Arsenous Acid.—It is employed usually as an alterative in doses of $\frac{1}{60}$ to $\frac{1}{20}$ grain.

Chromic Acid.

THIS is really an *oxide* of chromium, and resembles in point of faulty nomenclature the acid previously described. Crystals of the acid readily separate from a mixture of a concentrated solution of potassium bichromate and strong sulphuric acid. The acid should be added to the bichromate solution slowly and with constant stirring, to prevent undue heating. The crystals, after removal of the mother-liquor, are washed with nitric acid, allowed to drain, and preserved in small well-stoppered bottles. This equation illustrates the reaction:



The acid sulphate of potassium remains in solution.

Properties.—Chromic acid is in form of small crimson needles or columnar crystals, very deliquescent, and hence easily soluble in water. It attacks the skin, producing a caustic effect, wherefore it should always be handled with caution. It is best to use a glass spatula in manipulating it. Owing to the fact that it contains a large proportion

of oxygen it readily attacks and oxidizes a good many bodies, causing combustion or explosion. Glycerin, sweet spirit of nitre and strong alcohol are among the easily oxidizable bodies to which chromic oxide should not be added.

Uses.—Chromic acid is never used internally. It is valuable as an antiseptic and powerful caustic, and is used largely to remove warts and callous growths.

The Metals.

Physical Properties.—There is a wide range in the *densities* of the metals—lithium, sodium and potassium being the three which are lighter than water, while osmium is the heaviest, with iridium, platinum and gold following. The metals have all, more or less, a *metallic lustre*, their surfaces admitting of being highly polished. In *color* they range from the white of silver to the red of copper and yellow of gold. Bismuth is white, but has a faint pinkish hue.

Certain of the metals are *malleable*, a property which permits of their being extended by hammering or rolling. Gold is the most malleable of all the metals. Gold leaf is one of its attenuated forms.

Ductility is a property involving the principle of tenacity, or power of resisting tension, possessed by metals which admit of being drawn out into wire.

Volatility is probably possessed by all metals, but sufficiently high temperatures cannot be obtained to volatilize all. Mercury easily volatilizes at low temperature, but the highest degree of heat that we can to-day obtain only fuses or melts many of the metals.

Chemical Properties.—Besides the combinations before referred to, into which metals enter—i.e., with the non-metals, forming oxides and salts—they combine with other metals. The compounds formed by the union of metals among themselves are called *alloys*, excepting that when one of the metals is mercury, the union is called an *amalgam*. It is not positively known whether the metals in alloys combine chemically or not. One of the chief char-

acteristics of the chemical force is its property of totally changing the *physical* appearances and properties of the bodies acted upon. This is not observed in case of alloys: the physical properties are still those of metals, and there is no such obliteration of the identity of the constituents as ensues when a metal combines with, e.g., oxygen or sulphur, but on the other hand, the evolution of much heat in the formation of some alloys, notably those of potassium and sodium, would indicate that there is chemical action. The alterations of character are usually in hardness, melting-point, color, tenacity and specific gravity. The combination, if it is such, takes place when the metals are melted together. While some chemists prefer to look upon alloys as chemical combinations, yet admitting that the chemical force exerted is very feeble, others consider them mere mechanical mixtures or solutions of one metal in the other in the liquid state.

Some of the commoner alloys are:

Brass, consisting of copper and zinc;

Gun-metal, consisting of copper and tin;

Bronze, consisting of copper, zinc and tin;

German silver, consisting of copper, zinc and nickel;

Pewter or Britannia, consisting of copper, tin and antimony;

Shot, consisting of lead, with antimony or arsenic;

Gold coin, consisting of gold, copper and sometimes silver.

Cavities in teeth are usually filled with alloys specially prepared.

The process of *amalgamation* is often resorted to in the extraction of some of the metals. Gold, which principally occurs in the metallic state as gold dust, is obtained from auriferous sand by treatment with mercury. The *amalgam* formed is heated in a retort, whereby the mercury is re-

covered by distillation, the gold remaining as a residue in the retort. Silver is obtained from its poorer ores by the same method.

Unions of the Metals with the Halogens—Haloid Salts.—
CHLORIDES.—Chlorine combines with metals to form chlorides. The combinations take place in proportions determined by the atomicity of the metal, and the compounds formed may be looked upon as being derived from as many molecules of hydrochloric acid as the metal has bonds, by the replacement of the equivalent of hydrogen by a metal. This may be more easily comprehended by assuming hydrochloric acid to be the type of all chlorides.

HCl with potassium yields KCl, K replacing H;
 2HCl with calcium yields CaCl_2 , Ca replacing 2H;
 3HCl with gold yields AuCl_3 , Au replacing 3H;

Chlorides may be made in any of the following ways: 1. By acting on metals (antimony, copper, etc.) with elementary chlorine; or, 2, with hydrochloric acid (zinc, iron, etc.) 3. By acting on oxides (silver) with free chlorine; or, 4, with hydrochloric acid; and, 5, by dissolving carbonates or hydroxides in hydrochloric acid.

OXYCHLORIDES and **SULPHOCHLORIDES** are unions of chlorides with oxides and sulphides, respectively. Bismuth oxychloride may be viewed as being the chloride, BiCl_3 , with two of the chlorine atoms replaced by oxygen, BiOCl . Metallic chlorides combine in definite proportions with ammonia and organic bases. Diammonium-platinous chloride, $2\text{NH}_3.\text{PtCl}_2$, is such a compound, and is used occasionally in testing. Nearly all the chlorides are soluble. Notable exceptions are those of silver, lead and mercurous mercury.

BROMIDES.—Bromine combines with most of the metals, and produces compounds analogous in composition to the chlorides, which they resemble in most of their properties. Chlorine having stronger affinity for metals than bromine, replaces the latter in salts, setting it free. Strong sulphuric acid also liberates bromine from any bromide.

IODIDES.—Unions of iodine with the metals are called iodides. They are analogous in properties to the chlorides and bromides. Chlorine and strong sulphuric acid will liberate the iodine, which turns blue with starch paste.

CYANIDES are combinations closely related to the haloid salts, excepting that the radical is composed of two elements, carbon and nitrogen, together called *cyanogen* and written CN or Cy. Some cyanides are made by heating the metal in cyanogen gas or vaporized hydrocyanic acid. Cyanogen has its source chiefly in nitrogenous organic compounds, and a number of cyanides are made by heating the fixed alkalies with such bodies. A number of the cyanides unite with each other, producing the *double cyanides*. Ferricyanide and ferrocyanide of potassium are such double salts. K_4FeCy_6 is the union of $FeCy_2$ and $4KCy$ —ferrocyanide potassium. These double cyanides are important in analysis.

Oxygen Salts.—With oxygen the haloid salts produce a different class of salts, better known as belonging to the oxygen salts; the haloid salts, it should be noticed, contain no oxygen. The oxygen salts may be derived from the corresponding acid by replacing the hydrogen with a metal—potassium in this case:

(Potassium chloride, KCl , from HCl , hydrochloric acid.)

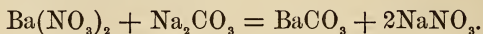
Potassium chlorite, $KClO_2$, from $HClO_2$, chlorous acid.

Potassium chlorate, $KClO_3$, from $HClO_3$, chloric acid.

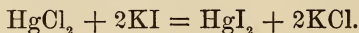
Iodine and bromine have corresponding salts, also obtained by replacing the hydrogen in corresponding acids. The names of the salts end as above.

In a previous chapter it was explained that oxides are conveniently divided into three classes,—viz., *acid*, *neutral* and *basic*,—and that the first and third are capable of uniting to form salts, but that in the majority of cases, *metallic salts* are obtained from *acids* by replacing their hydrogen by a metal, basic or neutral oxide or hydro oxide. It was also explained that acids are unions of acid oxides with water, the oxide of hydrogen, and hence are intermediate compounds between oxides and oxygen salts.

Another method of making salts is the interchange of the bases contained in soluble salts; sodium carbonate and barium nitrate will produce the insoluble carbonate of barium, which is precipitated, and soluble sodium nitrate, which remains in solution:



Insoluble salts containing halogen elements may also be made by this interchange method:



The mercuric chloride is decomposed, with the formation of insoluble mercuric iodide and soluble potassium chloride. All haloid salts and acids may therefore be looked upon as constructed after one type, hydrochloric acid being usually taken for the type, $\text{H}-\text{Cl}$. By replacing the chlorine with any of the other halogen elements the remaining halogen acids may be obtained, and by replacing the hydrogen in these all the various haloid salts are produced. The hydrogen of the acid or the metal in the salt is the *base* and the halogen element the *radical*. Now, oxygen salts may

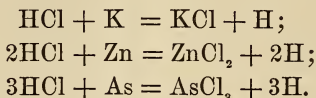
be regarded, as above stated, as compounds of acid oxides with basic oxides, *or* as analogous in composition to hydrochloric acid; that is, as being constructed after the hydrochloric acid type, containing a base and a radical, without reference to oxides. Thus the oxygen salts, sodium nitrate, NaNO_3 , and calcium carbonate, CaCO_3 , may be looked upon as constructed from HCl , the Na and Ca taking the place of hydrogen, and the *groups* of elements NO_3 and CO_3 the place of the chlorine. The NO_3 and CO_3 together as one radical discharge functions similar to those of chlorine, and like that element are capable of migrating unchanged and together from one compound to another.

It is a subject of discussion among chemists which of these views is the correct one. The former is a rational view and the latter a more convenient one. The latter is also known as the *binary theory of salts*.

Chemists agree that it is impossible to construct formulas capable of representing the exact arrangement of atoms in molecules; it is sufficient to be able to exhibit the combining power of the atoms of the several elements in compounds and the manner in which any group of atoms splits up into smaller subordinate groups under the influence of other compounds. To illustrate this sulphuric acid is a good example, as it splits up or decomposes in four ways, according to the reagent used: 1. With zinc it splits up into H_2 and SO_4 , the H_2 becoming liberated and the SO_4 combining with zinc to form its sulphate: $\text{Zn} + \text{H}_2\text{SO}_4 = \text{ZnSO}_4 + \text{H}_2$. 2. If lime, CaO , is added to sulphuric acid, it decomposes the latter into H_2O and SO_3 , the H_2O becoming free and the SO_3 uniting with the CaO to form calcium sulphate: $\text{CaO} + \text{H}_2\text{SO}_4 = \text{H}_2\text{O} + \text{CaSO}_4$. 3. Sulphuretted hydrogen is formed when sulphide of iron, FeS , is added to sulphuric

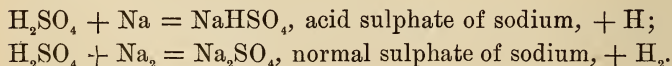
acid; H_2SO_4 in this case decomposes into H_2S and O_4 .
 $\text{H}_2\text{SO}_4 + \text{FeS} = \text{H}_2\text{S} + \text{FeSO}_4$. 4. With barium dioxide, BaO_2 , sulphuric acid splits up into H_2O_2 and SO_2 ; $\text{H}_2\text{SO}_4 + \text{BaO}_2 = \text{H}_2\text{O}_2 + \text{BaSO}_4$.

Normal, Acid and Double Salts.—Monobasic acids produce *normal salts*, by the replacement of their hydrogen by a metal. The metal may take the place of the hydrogen in one, two or three molecules of hydrochloric acid, for instance:



A normal salt is therefore one derived from an acid by replacing *all* the hydrogen in it by a metal: CaCO_3 , BaSO_4 .

Acids which are *bibasic* or *dibasic*—that is, which have two replaceable hydrogen atoms—may produce two kinds of salts, *normal* and *acid* salts. The normal, as in case of salts derived from monobasic acids, have an equivalent of the metal in place of *all* the hydrogen, while the acid salt still contains some hydrogen, only part of it having been replaced in the acid by a metal:



An acid salt, therefore, is a salt containing a metal *and* hydrogen in its base ($\text{NaH}(\text{CO}_3)$, $(\text{KH})(\text{C}_4\text{H}_4\text{O}_6)$).

If the hydrogen in a *dibasic* or *tribasic* acid be replaced by *two* metals a *double salt* results: Rochelle salt, double tartrate of sodium and potassium, $\text{KNaC}_4\text{H}_4\text{O}_6$, is derived from tartaric acid, $\text{H}_2\text{C}_4\text{H}_4\text{O}_6$, by replacing one of its hydrogen atoms with potassium, the other with sodium.

Bibasic acids may therefore produce *normal* or neutral

salts, as they are sometimes called, *acid* or *bisalts*, and *double salts*.

Tribasic acids may produce as many kinds of salts as the bibasic, and in addition tertiary salts, having three metals in the base.

Basic Salts or Oxysalts.—Normal salts give rise to two other classes of salts, the *basic salts* or *oxysalts*, and *anhydro salts*. The composition of these will be better understood by looking upon salts again as compounds of acid oxides with basic oxides, as stated above. (The terms basic oxide and basic salt should not be confounded by the student.)

Normal sodium sulphate, Na_2SO_4 , equals Na_2O and SO_3 ;
 Normal bismuth carbonate, $\text{Bi}_2(\text{CO}_3)_3$, equals Bi_2O_3 and 3CO_2 ;

Normal ferric sulphate, $\text{Fe}_2(\text{SO}_4)_3$, equals Fe_2O_3 and 3SO_3 .

In these it will be seen that there are as many acid oxide molecules as there are oxygen atoms in the basic oxide. When the proportion of the acid oxide is less, the salt is termed a *basic salt*:

Basic bismuth carbonate, $\text{Bi}_2\text{O}(\text{CO}_3)_2$, is Bi_2O_3 and 2CO_2 ;

Basic bismuth carbonate, $\text{Bi}_2\text{O}_2\text{CO}_3$, is Bi_2O_3 and CO_2 ;

Basic ferric sulphate, $\text{Fe}_2\text{O}_2\text{SO}_4$, is Fe_2O_3 and SO_3 ;

Basic ferric sulphate, $\text{Fe}_2\text{O}(\text{SO}_4)_2$, is Fe_2O_3 and 2SO_3 .

It will be noticed that the acid oxides are less in number than the number of atoms of oxygen in the basic oxide. There is an analogy of these basic oxysalts with the oxychlorides, oxybromides, etc.: thus the basic bismuth chloride has the formula BiOCl or $\text{Bi}_2\text{O}_2\text{Cl}_2$, the group of elements CO_3 in the basic carbonate acting as the single element does in the basic chloride. Some normal salts absorb moisture from the air; the moisture being water, H_2O , there

is consequently a greater proportion of base present than of acid, and a basic salt results: $4\text{FeSO}_4 + \text{O}_2 = \text{Fe}_2(\text{SO}_4)_3$, normal ferric sulphate, and $\text{Fe}_2\text{O}_3\text{SO}_4$, basic ferric sulphate. The FeSO_4 is a ferrous salt and by the absorption of oxygen becomes *oxidized* to the ferric salt.

The second class of salts of the normal kind are those which contain an excess, not of the basic oxide, as in the above, but of the acid oxide. Thus:

Sodium anhydrosulphate, $\text{Na}_2\text{S}_2\text{O}_7$, is Na_2O and 2SO_3 ;

Potassium anhydrochromate, $\text{K}_2\text{Cr}_2\text{O}_7$, is K_2O and 2CrO_3 .

There are two acid oxides to one basic oxide in each of these salts. They are sometimes called *acid salts* because of the excess of acid oxide, and sometimes *bi salts* for the same reason, and because in the above cases they have *two* molecules of the acid oxides to designate. The prefix *anhydro* should designate, when added to the name of a normal salt, an addition to the latter of a molecule of the anhydrous acid. By anhydrous acid is meant that which is left after all the hydrogen and enough oxygen to form water have been removed from an acid; the *anhydride* of sulphuric acid is SO_3 : $\text{H}_2\text{SO}_4 = \text{H}_2\text{O} + \text{SO}_3$. SO_3 added to Na_2SO_4 gives $\text{Na}_2\text{SO}_4\text{SO}_3$, or $\text{Na}_2\text{S}_2\text{O}_7$. So CrO_3 , added to K_2CrO_4 , gives the so-called bichromate of potassium, $\text{K}_2\text{Cr}_2\text{O}_7$.

Kinds of Formulas.—Formulas are usually collections of symbols employed to express the composition of compounds. The simplest is the *empirical* formula, which expresses merely the *proportions* of elements in compounds. Benzene, C_6H_6 , contains six molecules of hydrogen and as many of carbon; hence its empirical formula is CH .

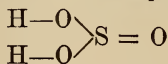
A **MOLECULAR FORMULA** expresses, in addition to this numerical relation, the actual number of atoms in a mole-

cule. C_6H_6 is the molecular formula of benzene, $NaCl$ for common salt, chloride of sodium, etc.

A RATIONAL FORMULA shows at a glance the disposition of the compound toward reagents or other compounds having chemical affinity. The rational formula for sulphuric acid is H_2SO_4 , which is to show that the H_2 or the SO_4 may be replaced; of hypophosphorous acid $H(PH_2O_2)$, meaning that only one hydrogen atom may be replaced, and that the (PH_2O_2) together form a group which behaves chemically as a single element would. H_3PO_2 would not be a rational formula for hypophosphorous acid.

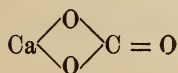
A GRAPHIC FORMULA indicates the equivalent or combining values of atoms, and the manner in which these are satisfied by combination. This is done by arranging the symbols in diagrams in such a way that each atom is connected with others by a number of lines or bonds corresponding to its atomicity or degree of equivalence, a monad having one bond; a diad, two; a triad, three, etc. For example:

Sulphurous acid, H_2SO_3 .



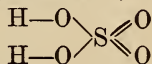
Here the S has four bonds, H one and O two.

Calcium carbonate,
 $CaCO_3$.



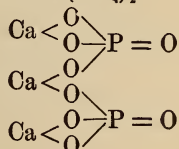
Ca has two, carbon four and oxygen two bonds.

Sulphuric acid, H_2SO_4 .



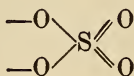
In this the S has six bonds.

Calcium phosphate,
 $Ca_3(PO_4)_2$.



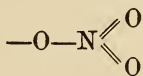
Ca has two, O two, and P five bonds.

Graphic formulæ are also called *structural* or *constitutional*. They do not represent the actual arrangement of the atoms in the molecule; chemists are ignorant of how atoms are arranged, and if they knew they could not represent the arrangement on a plane surface. It will be seen from the above diagrams that each atom has its bonds satisfied by union with a bond belonging to some other atom. The radical of a salt or acid is often spoken of as an *unsaturated group of elements*. The SO_4 is the sulphuric radical and has two free bonds, whose places in sulphuric acid are occupied by hydrogen:



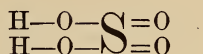
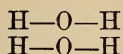
The two free bonds may be saturated with two atoms of a monad metal, or with one atom of diad metal. The (SO_4) acts as an element with two bonds would.

In the same way the nitric radical (NO_3) performs all the functions of a single atom with one bond:



In writing graphic formulas for salt containing oxygen, the water type is usually employed. $\text{H}_2\text{O} = \text{H—O—H}$, the left hydrogen in case of salts being replaced by metals, and the right by a single element or a group of elements. The number of bonds which the base has determines the number of molecules of water to be taken.

Sulphuric acid has two bonds in the base, H_2 , hence the graphic formula must be constructed after two molecules of water:



It will be seen that the two right-hand hydrogen atoms are replaced by SO_2 .

Written after the hydrochloric-acid type, water is derived from HCl by replacing the Cl with OH , and, *vice versa*, the HCl from H_2O by replacing the OH with Cl .

The student is urged to read again the opening lessons under the Mineral Acids, to which the above is supplemental.

The Alkalies.

IN the beginning of the course the student learned that the elements are divided, for convenience of study, into divisions and subdivisions, and that the two primary divisions represent metals and non-metals respectively. All of the important non-metals have now been studied individually and in their important combinations with each other. The acids, the last subject of study, may be looked upon as belonging to the latter—i.e., to combinations containing *only* non-metals, *if* hydrogen is considered a non-metal, which is disputed by some on the seemingly valid ground that in its chemical behavior it resembles the metals more closely than the non-metals. The student will now go furthur and confine his attention to the *metals* and their combinations with the non-metals, which combinations are termed *salts*. Salts are usually unions of metals with some *combination* of non-metals, the combination usually, *not always*, being an acid. An acid may be viewed, therefore, in the additional sense of being a prepared condition of one or more non-metals for combining with a metal or metals. (See introductory chapter to the study of the Inorganic Acids.)

The first convenient subdivision of the metals is the *alkalies*, or alkali metals—sodium, potassium, lithium, cæsium and rubidium. The latter two are of no importance to the pharmacist. The hypothetical ammonium is usually included in this division.

They are called alkalies because they were formerly, and still are, to an extent, obtained from the ashes of a plant called *glasswort* (which ashes yielded not only soda

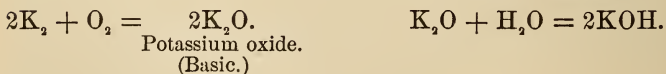
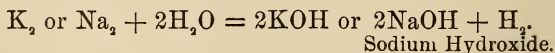
and potassa, but which were also employed in the manufacture of glass). The name of the ashes in Arabic is *gali*; in French and Spanish, *cali*. The Arabic prefix *al* is equivalent to our English article *the*. The Arabic word *galaj* signifies "to roast in a pan," and originated probably from the manner in which the ashes were collected. Soda and potassa were obtained by a process of lixiviation—i.e., by treating the ashes with water, which dissolved out both substances and yielded them up again upon evaporation. This was the original method of obtaining soda and potassa.

Properties.—The alkali metals resemble each other in many of their physical properties. There is one dissimilarity, however, that is marked. Sodium, potassium and lithium and their salts are fixed, whereas ammonium and its salts are volatile. This distinction is brought to mind by the terms "fixed alkalies" for the former and "volatile alkali" for the latter. *Fixed* means remaining permanent upon the application of heat, and *volatile*, dissipating or entering the gaseous state in contact with heat. In *composition* there is also a difference between the fixed alkalies and the volatile alkali; each of the former consist of a single element, K, Na, and Li, while the latter is *compound*, consisting of the two elements hydrogen and nitrogen, NH_3 , but the NH_4 is purely hypothetical, and used for convenience— NH_3 is the base. This *may not* be a distinction; that which is termed ammonium we have succeeded in decomposing into the two elements hydrogen and nitrogen, while sodium, potassium and lithium have resisted up to the present all attempts at further subdivision, and hence are called elements.

The bodies which we now call elements may be compounds whose constituents have more chemism or chemi-

cal affinity for each other than for other bodies. It may be that there are bodies which, if our elements or some of them are compounds, have greater affinity for one or the other of the constituents, but if there are, they have never been brought in contact or condition to exert their greater power of chemical affinity. Whenever the latter happens a new element or elements will be discovered. The annals of chemical history verify this, as now and then a body which had been supposed to be simple is found to be compound.

The alkaline hydroxides resemble each other in the following common properties: They (1) turn red litmus-paper blue, and (2) colorless phenol-phthalein solution to a fuchsine red, (3) neutralize acids, (4) form salts with acids, which are neutral or acid according to the proportion of the acid, and sometimes alkaline; (5) saponify fats—that is, form soap with fatty bodies, and (6) cauterize the skin. The metals of the alkalies are all soft (easily cut with a knife), all lighter than water, fuse very readily at low temperatures and volatilize at high temperatures, decompose water with the evolution of hydrogen, combine with great energy with oxygen, forming strongly basic oxides, which dissolve easily in water, producing alkaline hydroxides, which are powerfully caustic:



They all form alums and produce *basic oxides* with oxygen—that is, they will unite with acids to make salts in which they will act as *base*. *Acid oxides*, such as SO_2 ,

P_2O_5 , CO_2 , etc., constitute the acid portion of a salt; thus, sulphite of sodium, Na_2SO_3 , is composed of:

The *base*..... K_2O , and

The radical or acid..... SO_2

Making the salt..... K_2SO_3

Attention is again directed to the synonymy of the terms *base* and *basyulous radical*, and *radical* and *acidulous radical*. In the above salt the K_2O may be termed simply the *base* or the *basyulous radical*, and the SO_2 the *radical* or the *acidulous radical*. The latter form of nomenclature assumes that a salt is composed of two *radicals* or *roots*, one a base and the other an acid. The former mode is preferable for its brevity—i.e., base and radical. The metals of the alkalies form only one chloride each, their salts are all colorless and soluble; and all but ammonium give characteristic colors to colorless flames. Sodium colors the Bunsen-burner flame intensely yellow, potassium gives it a distinct purple hue if viewed through blue glass, and lithium produces a bright crimson color. Cæsium means blue, and rubidium, red.

Alums.

It is a peculiar fact that all the metals of the alkalies form alums, double salts of great interest.

An alum may be defined to be a double salt formed by the union of the sulphate of an alkali metal with the sulphate of a pseudo-triad, and having 24 molecules of water of crystallization.

For illustration: Potassa-alum has the formula $K_2Al_2(SO_4)_4 \cdot 24H_2O$, which shows that theoretically it is composed of:

Potassium sulphate.....	K_2SO_4
Aluminum sulphate (pseudo-triad).....	$K_2Al_2(SO_4)_3$
	$K_2Al_2(SO_4)_4$

In crystallizing it takes up 24 molecules of water.

The following alums may be resolved as the potassa alum:

Iron alum, $K_2Fe_2(SO_4)_4 \cdot 24H_2O$, containing potassium.

Iron alum, $(NH_4)_2Fe_2(SO_4)_4 \cdot 24H_2O$, containing ammonium.

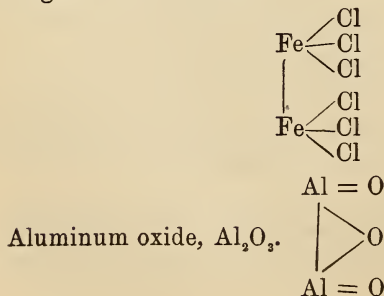
Chrome alum, $(NH_4)_2Cr_2(SO_4)_4 \cdot 24H_2O$.

Cæsium alum, $Cs_2Al_2(SO_4)_4 \cdot 24H_2O$, etc.

Iron, chromium, aluminum, and other metals which have four bonds, of which only three are active, may be called pseudo-triads, for want of a better term. The proportions in which they unite with other elements and the formulas obtained for some of their salts in ultimate analysis tend to show that they have in some cases three active bonds, but never in the case of single atoms; *two* atoms always have together *six* bonds.

Graphic formulæ serve to illustrate this; for instance:

Ferric chloride, Fe_2Cl_6 , not $FeCl_3$, may have its atoms arranged in this manner:



The atoms of iron and aluminum have each four bonds, but two bonds, one from each atom, neutralize or combines each other, so to speak, so that *two atoms together have six bonds*. It will not be inappropriate, therefore, to call these metals false or pseudo-triads. (Recent investigations made with aluminum make it doubtful whether it is a pseudo-triad. The vapor density of some of its combinations would seem to indicate that it is a true triad.) The alkali metals have each only one bond and are probably the only monad metals entering into the composition of the alums.

Ammonium is only a hypothetical metal, and is classed with the alkali metals because of the similarity it has in its chemical behavior to the metals of that group.

Potassium Salts.

THE salts of potassium are very numerous, and are among the most important salts used in pharmacy. They are colorless when crystallized, and white when powdered or granulated. Excepting the bitartrate, sulphate and chlorate, they are very soluble in water; those mentioned are soluble to a less extent. Owing to their extreme solubility there is no good test in which precipitation may be employed as a means of determining their presence or identity. The greater number of the salts of potassium are insoluble in alcohol.

Sources.—Formerly the product obtained by lixiviating *wood-ashes* was the only source, but to-day the chief sources are the products of the Stassfurt mines in Germany, which consist of a variety of potassium salts, among which the chloride predominates. The salts are converted into sulphate and this into carbonate, from which most of the pharmacopœial salts are made.

Argols is the chief source of cream of tartar, bitartrate of potassium. This salt exists naturally in the juice of grapes, but deposits upon the sides of the casks when fermentation sets in, which changes the sugar into alcohol, and precipitates the salts, which are (as mentioned above) insoluble in alcohol.

Suint, or the soapy or fatty substance obtained from sheep's wool, yields a number of potassium salts, which are obtained by washing the wool, filtering the washings, concentrating and crystallizing.

Some potassium salts are obtained in the extraction of sugar from beet-roots. The mother-liquor left after the third or fourth crystallization of sugar, and from which no more crystals can be obtained, and which constitutes beet-root molasses, contains enough potassium salts to make it profitable to extract them.

Calcutta nitre, nitrate of potassium, is also found in Nature. See below.

Tests for Potassium Salts.—The colorless flame of the Bunsen-burner is colored violet or purple, best seen through blue glass.

When a strong solution of a potassium salt is added to excess of a concentrated solution of tartaric acid, a white crystalline precipitate slowly forms. The precipitation may be hastened by the addition of alcohol; the precipitate is cream of tartar, which is the least soluble of the potassium salts.

The addition of a solution of platinic chloride containing a little alcohol and hydrochloric acid produces in strong solutions of potassium salts a yellow precipitate consisting of the double chloride of potassium and platinum, $\text{PtCl}_4 + 2\text{KCl}$.

In the collective study of the potassium salts it is best to begin with the natural products and follow the preparation of these into the various pharmacopœial salts. The alphabetical arrangement of the U. S. P. is best for *classification*; in studying the formation of these salts it is simpler to follow their derivation.

Impurities.—By noting the impurities in the *sources* it will not be difficult to remember the possible impurities in the various salts, as it rarely happens that contamination takes place during their preparation. The chief impurities in the several sources which may be looked for in the various salts by the respective tests are:

CHLORIDES.—Precipitated by AgNO_3 .

SULPHATES.—Precipitated by solution BaCl_2 .

CARBONATES.—In some, effervescence with acids.

ALKALINE EARTHS.—The neutralized solution of a salt gives cloudiness or white precipitate with solution Na_2CO_3 .

ORGANIC IMPURITIES.—Strong sulphuric acid will become blackened if fragments of a salt containing organic matter are sprinkled upon its surface.

SILICA.—In the presence of silica a solution acidulated with nitric acid and evaporated to dryness yields a residue which is not wholly soluble in water.

METALS.—A solution of a salt acidulated with nitric acid gives a black color or precipitate with sulphuretted hydrogen or ammonium sulphide in the presence of metals.

The foregoing are general impurities; specific contaminations will receive mention in appropriate places.

Potassium Carbonate. *Potassii Carbonas.* K_2CO_3 .—From this salt is obtained or prepared the majority of the pharmacopœial preparations of potassium.

Potassium carbonate is obtained from the products of

the Stassfurt mines by a process of conversion, or from wood-ashes by repeated crystallizations. It is a white crystalline or granular powder, odorless and strongly alkaline, deliquescent, and, in large doses, caustic. Lemon juice or diluted vinegar is an appropriate antidote. *Dose*, 5 to 15 grains (0.33 to 1.00 gm.).

Potassium Bicarbonate. *Potassii Bicarbonas.* KHCO_3 .—Obtained by saturating a solution of carbonate of potassium with carbonic acid gas, evaporating at a low temperature (70°C.) and crystallizing—e.g., $\text{K}_2\text{CO}_3 + \text{H}_2\text{O} + \text{CO}_2 = 2\text{KHCO}_3$. It occurs in transparent monoclinic prisms, permanent in dry air, odorless, of slightly alkaline reaction, decomposed by heat or boiling into carbonate and CO_2 —e.g., $2\text{KHCO}_3 = \text{K}_2\text{CO}_3 + \text{H}_2\text{O} + \text{CO}_2$. This salt is usually quite pure, and is the starting-point in the manufacture of a number of other salts. It is also valuable because of the large proportion of carbonic acid gas it can be made to yield. It is milder than the carbonate; chemically, it is an acid salt, because it contains hydrogen in the base.

Potassium Sulphite. *Potassii Sulphis.* K_2SO_3 .—If sulphurous acid gas is passed through solution of potassium carbonate until all of the carbonic acid is expelled, potassium sulphite will be formed—e.g., $\text{K}_2\text{CO}_3 + \text{SO}_2 = \text{K}_2\text{SO}_3 + \text{CO}_2$. A little sulphate is usually formed, but the U. S. P. of 1880 required at least 90 per cent of sulphite.

This salt is in white, opaque crystals, or in crystalline powder slightly deliquescent, odorless, and of a bitter, sulphurous taste. It is possessed of valuable anti-fermentative properties, retarding or checking various kinds of fermentation and destroying lower forms of organic life.

Dose ranges from 10 to 40 grains (0.66 to 2.66 gms.).
Formula, $K_2SO_3 \cdot 2H_2O$. It is no longer official.

Potassium Acetate. *Potassii Acetas.* $KC_2H_3O_2$.—Made by neutralizing acetic acid with potassium bicarbonate.
 $KHCO_3 + HC_2H_3O_2 = KC_2H_3O_2 + H_2O + CO_2$. The solution is carefully evaporated in a porcelain dish, avoiding contact with iron (iron gives red color by forming the red acetate of iron). This salt may be made from the carbonate, but the bicarbonate should be preferred because of its greater purity. Organic matter may be present from the acetic acid, which is of organic origin. It is in white, foliaceous, satiny masses, or in white powder, very deliquescent, should be odorless, but often has faint odor of acetic acid, neutral or faintly alkaline reaction; soluble in alcohol. It is a diuretic. Dose, 1 to 4 gms.

Potassium Citrate. *Potassii Citras.* $K_3C_6H_5O_7$.—Made by adding crystals of citric acid to a solution of bicarbonate of potassium until neutralized or until effervescence ceases, evaporating and crystallizing: $3KHCO_3 + H_3C_6H_5O_7 = K_3C_6H_5O_7 + 3H_2O + 3CO_2$. It is a white granular powder, deliquescent, odorless, slightly alkaline taste, and neutral or faintly alkaline reaction. In addition to general impurities it may contain *tartrate*, which is detected by adding acetic acid to a concentrated solution of the salt—in its presence a white crystalline precipitate of the acid tartrate will deposit. It is diaphoretic in doses of 0.66 to 1.33 gms.

Solution Potassium Citrate. *Liquor Potassii Citratis.* Made in a manner similar to that employed for the preparation of the salt, without evaporating the solution. It is a clear, colorless liquid, odorless, containing about 9 per cent of

the salt, and should be prepared as needed. As refrigerant and diaphoretic, the dose is 3 to 5 cc.

Mixture Potassium Citrate.—**NEUTRAL MIXTURE.**—This mixture resembles the foregoing in all respects, excepting, perhaps, that the use of lemon juice in place of citric acid renders it more agreeable in taste. Properties and dose the same as those of the solution. This mixture has been dismissed from the 1890 U. S. P.

Effervescent Potassium Citrate. *Potassii Citras Effervescens.*—This preparation has been added to the Pharmacopœia in its last revision. Potassium bicarbonate and citric acid are intimately mixed in a warm mortar, and the resulting paste rapidly dried at or below 120° C., and then reduced to any desired degree of fineness by powdering. The bicarbonate and the acid are incompatible, and in the presence of moisture or water reaction ensues, with the liberation of carbonic acid gas: $3\text{KHCO}_3 + \text{H}_3\text{C}_6\text{H}_5\text{O}_7 = \text{K}_3\text{C}_6\text{H}_5\text{O}_7 + 3\text{H}_2\text{O} + 3\text{CO}_2$. The preparation should be kept protected from dampness.

Solution of Arsenite of Potassium.—Fowler's solution (see under Arsenic).

Sulphurated Potassa. *Potassa Sulphurata.*—This preparation is made by heating together carbonate of potassium and sublimed sulphur until the mixture is completely melted, pouring on a stone slab, and breaking into pieces when cooled. *Liver of Sulphur* is a name sometimes given to it. It is not a definite chemical compound, but consists of varying quantities of sulphide and hyposulphite of potassium. The U. S. P. requires the presence of at least 12.85 per cent of sulphur combined with potassium to form sulphide. The pieces are of a liver-brown to a brownish-yellow color, of a faint, disagreeable odor and

repulsive taste. Nearly wholly soluble in water, only partly in alcohol. Though sometimes given internally, it is usually employed externally in skin diseases.

Potassium Hypophosphite. *Potassii Hypophosphis.* KH_2PO_2 .
—Made by acting on the potassium carbonate with calcium hypophosphite, filtering the solution, evaporating and granulating: $\text{K}_2\text{CO}_3 + \text{Ca}(\text{H}_2\text{PO}_2)_2 = 2\text{KH}_2\text{PO}_2 + \text{CaCO}_3$. The salt occurs usually as a white granular powder, is very deliquescent, odorless, of sharp, bitter taste, soluble in alcohol. The U. S. P. guards against the presence of phosphate, by directing that not more than a slight cloudiness should appear on mixing an aqueous solution of the salt with test solution of magnesium. Potassium hypophosphite is one of the ingredients of syrup of hypophosphites. Dose is from 5 to 15 grains (0.33 to 1.00 Gm.).

Potassa. KOH .—Potassa is the starting-point in the manufacture of another series of the potassium salts. It is made by boiling together potassium carbonate, or, preferably, the bicarbonate, and slaked lime, setting the mixture aside until clear and removing the clear solution by means of a glass syphon, and evaporating in a clean iron, or better, silver vessel until a fluid of an oily consistence is obtained. The hot *caustic potassa* is poured into moulds, producing the sticks of the market: $\text{Ca}(\text{OH})_2 + 2\text{KHCO}_3 = \text{CaCO}_3 + 2\text{KOH} + \text{H}_2\text{O} + \text{CO}_2$.

Potassa by Alcohol is potassa purified by dissolving in alcohol—the alcohol dissolves only the potassa, leaving the impurities behind, which may be removed by filtration. Caustic potassa is in form of sticks usually, hard, very deliquescent, odorless or having a faint odor of lye, very caustic and strongly alkaline. It is principally used as a caustic. *Antidotes* are diluted vinegar, or citric acid solu-

tion, or lemon juice, followed with a bland oil to correct its irritant effects. It should not be handled with the fingers. The commercial product contains from 20 to 40 per cent moisture.

Solution of Potassa. *Liquor Potassæ*.—Made as the above without evaporating, or by dissolving caustic potassa in water. It is an aqueous solution containing about five per cent of KOH, clear, colorless, odorless, very acrid and caustic taste. All that has been said about potassa applies to the solution. It is given internally as an antacid in doses of 0.5 to 1.00 cc. diluted.

Potassa with Lime. *Potassa cum Calce*. KOH and CaO.—Made by rubbing together equal weights of potassa and dry lime. This preparation is a grayish-white powder, very alkaline in appearance, and deliquescent. Sometimes it is found in sticks similar to the caustic potassa sticks. Used externally as caustic, but is milder and slower in its operation than potassa alone.

Potassium Permanganate. *Potassii Permanganas*. $K_2Mn_2O_8$ or $KMnO_4$.—Made by heating together caustic potassa, chlorate of potassium and dioxide of manganese, and boiling the resulting mass in water to change the formed manganate into the permanganate, which is obtained from the resulting solution by careful evaporation: (1) $6KOH + 3MnO_2 + KClO_3 = 3K_2MnO_4$ (potassium manganate) + $KCl + 3H_2O$; (2) $3K_2MnO_4 + 2H_2O = 2KMnO_4$ (potassium permanganate) + $MnO_2 + 4KOH$. The potassium hydroxide is usually neutralized with dilute H_2SO_4 , because not all of the manganate is oxidized in presence of excess of potassa. The salt occurs as deep purple, violet or nearly black, needle-shaped prisms of a metallic lustre,

permanent, without odor, sweet, astringent taste, and is easily decomposed by alcohol.

Permanganate of potassium is one of the most powerful of our oxidizing agents. The U. S. P. directs that "it should not be triturated nor combined in solution with organic or readily oxidizable substances." It parts easily with its oxygen, this combining with organic matter. Chemically it is used as a volumetric test and oxidizer, but its solution should not be kept too long, as light and air decompose it. It is an excellent disinfectant.

Potassium Iodide. *Potassii Iodidum.* KI.—This, the most important of the potassium salts, is prepared by adding a slight excess of iodine to an aqueous solution of potassa, evaporating to dryness and mixing the product with charcoal, which is then heated to dull redness. The resulting mass contains iodide of potassium, which may be dissolved out and purified by crystallization, e.g.: $6\text{KOH} + 3\text{I}_2 = 5\text{KI} + \text{KIO}_3 + 3\text{H}_2\text{O}$. The charcoal removes the oxygen from the iodate (KIO_3): $2\text{KIO}_3 + 3\text{C}_2 = 2\text{KI} + 6\text{CO}$. This salt is sometimes made from ferrous iodide (see process under Bromide). Iodide of potassium is in form of colorless, translucent cubical crystals, somewhat deliquescent, peculiar faint odor, saline, bitter taste, soluble in 18 parts alcohol.

When pure it is neutral, but the product of the market has a slightly alkaline reaction. This is desirable, since a slight alkalinity prevents the separation of the iodine. The U. S. P. allows the presence of 0.5 per cent of chloride or bromide. This iodide, like other soluble ones, precipitates alkaloids, hence is incompatible with these. It is official also in form of ointment.

Potassium Bromide. *Potassii Bromidum.* KBr.—This salt is

made by two methods. One is identical with that employed in making the iodide, bromine being used in place of iodine. By substituting the symbol Br for I, the student will have the reactions for the process. The other method employs ferrous bromide and potassium carbonate. The ferrous bromide is made by acting on iron with bromine : $\text{Fe}_2 + 2\text{Br}_2 = 2\text{FeBr}_2$; $\text{FeBr}_2 + \text{K}_2\text{CO}_3 = \text{FeCO}_3 + 2\text{KBr}$. The iron becomes an insoluble carbonate, and precipitates as such, while the KBr remains in solution, from which it is obtained in crystals on evaporating the solution. The fact that this salt was formerly imported, but is now made in this country on a scale large enough to supply the demands of other countries besides our own, is good evidence of the progress of chemical industry in this country. Bromide of potassium is in form of translucent cubical crystals, permanent, odorless, pungent saline taste and neutral reaction. The commercial product is usually in white opaque crystals, having a slightly alkaline reaction to prevent the liberation of bromine. The U. S. P. allows the presence of 3 per cent of chloride. The Pharmacopœia gives tests for the presence of bromate and iodide and the usual impurities. Bromide of potassium is a nervine; its dose is from 0.33 to 1.33 gms.

Potassium Ferrocyanide. *Potassii Ferrocyanidum.* $\text{K}_4\text{FeCy}_6 \cdot 3\text{H}_2\text{O}$.—YELLOW PRUSSIAN OF POTASH.—Made by heating nitrogenous animal matter with crude potash, lixiviating the mass and treating with freshly-made ferrous carbonate and crystallizing. It occurs in large lemon-yellow translucent prisms or tablets, somewhat effervescent in air, odorless. Unlike most other salts containing cyanogen (CN), it is not poisonous when pure. It is recognized by the U. S. P., not for its medicinal properties, for it is seldom

employed as a remedy, but for its reliability as a test for ferric salts, with which it gives a deep blue color and precipitate.

Potassium Cyanide. *Potassii Cyanidum.* KCN.—By heating a mixture of potassium carbonate and ferrocyanide of iron to redness, potassium cyanide is obtained. Of the several qualities found in the market only the granular white variety should be employed medicinally. It is deliquescent, odorless, of a bitter-almond taste, strongly alkaline reaction and very poisonous. Its toxic action is the same as that of hydrocyanic acid. Dose is $\frac{1}{16}$ to $\frac{1}{8}$ grain (.004 to .008 gm.).

Potassium Bichromate. *Potassii Bichromas.* $K_2Cr_2O_7$.—Made by adding strong sulphuric acid to a solution of potassium chromate: $2K_2CrO_4 + H_2SO_4 = K_2Cr_2O_7 + K_2SO_4 + H_2O$. The difference in the solubility of the two salts affords a means of separating them by fractional crystallization. The salt is in form of large orange-red, four-sided prisms; permanent, odorless, metallic taste. It is used almost exclusively in form of test-solution. In large doses it is poisonous; chalk, magnesia or soap are appropriate antidotes.

Potassium Chlorate. *Potassii Chloras.* $KClO_3$.—The salt is now made by reacting with calcium chlorate upon potassium chloride: $Ca(ClO_3)_2 + 2KCl = CaCl_2 + 2KClO_3$, and crystallizing. Occurs in colorless plates of a pearly lustre, permanent, odorless, characteristic saline taste, neutral reaction *explosive*. Great care should be observed in handling potassium chlorate; it should not be triturated with bodies that are easily oxidized. The salt readily parts with its oxygen when brought into contact with bodies having an affinity for that element. If heated it loses its oxygen:

$2\text{KClO}_3 + \text{heat} = 2\text{KCl} + 3\text{O}_2$. This fact is taken advantage of in the preparation of oxygen for chemical and medicinal uses. Internally the salt is given in doses of from 0.20 to 0.66 gm.

Troches of Potassium Chlorate each contain 5 grains of the salt. It should be noted that all of the foregoing salts are derivatives of potassium carbonates.

Potassium Bitartrate. *Potassii Bitartras.* $\text{KHC}_4\text{H}_4\text{O}_6$.—Cream of tartar is obtained from grape-juice, in which it occurs naturally. Being insoluble in alcohol, it separates as soon as the sugar in the juice changes by fermentation into alcohol. The deposit thus formed on the interior of wine casks is known as argols, and when purified by passing its solution through animal charcoal and evaporating, constitutes cream of tartar. The salt is usually in form of white, gritty powder, odorless, of pleasant acid taste, and soluble in 210 parts of water, being the least soluble of the potassium salts.

Calcium tartrate is also present in grape-juice, and the U. S. P. allows the presence of 6 per cent in cream of tartar. It is refrigerant in doses of 4 gms. Tartaric acid and most of the tartrates are made from it. Compound-jalap powder contains it as one of its ingredients.

Potassium Tartrate. *Potassii Tartras*, 1880. $\text{K}_2\text{C}_4\text{H}_4\text{O}_6$.—This salt is no longer official. It is made by reacting on cream of tartar with potassium carbonate: $2\text{KHC}_4\text{H}_4\text{O}_6 + \text{K}_2\text{CO}_3 = 2\text{K}_2\text{C}_4\text{H}_4\text{O}_6 + \text{H}_2\text{O} + \text{CO}_2$. It occurs as white powder, somewhat deliquescent, odorless, saline, bitter taste; neutral reaction. In doses of 8 to 30 Gm. it is purgative.

Potassium and Sodium Tartrate. *Potassii et Sodii Tartras.* $\text{KNaC}_4\text{H}_4\text{O}_6$.—ROCHELLE SALT.—The process for making

this double salt is the same as that employed for making the neutral tartrate, excepting that sodium carbonate is used in place of potassium carbonate: $2\text{KHC}_4\text{H}_4\text{O}_6 + \text{Na}_2\text{CO}_3 = 2\text{KNaC}_4\text{H}_4\text{O}_6 + \text{H}_2\text{O} + \text{CO}_2$. Rochelle salt is in form of white powder, becoming hard in moist air, odorless, mildly saline taste, neutral reaction. It is a purgative in doses of 15 to 30 Gm. Seidlitz powders contain it as one of their ingredients.

This salt, which contains two metals in the base, is a good illustration of a double salt; it may be looked upon as being thus constituted: $2\text{KNaC}_4\text{H}_4\text{O}_6 = \text{K}_2\text{C}_4\text{H}_4\text{O}_6 + \text{Na}_2\text{C}_4\text{H}_4\text{O}_6$.

Potassium Nitrate. *Potassii Nitras.* KNO_3 .—Nitre or salt-petre, or Calcutta nitre, occurs as an efflorescence upon the soil in India, and is purified by repeated crystallization. The salt is in form of colorless, transparent, six-sided prisms, or as a crystalline powder, permanent in the air, odorless, cooling, pungent taste, neutral reaction. It is diaphoretic and diuretic in doses of 0.66 Gm. In the arts it is used largely in making gunpowder. The cheaper sodium nitrate is not suitable for this purpose on account of its deliquescence. The nitrate is also official in form of a *paper*, made by immersing unsized paper in a strong solution of the nitrate and drying. The product is the *Charta Potassii Nitratis*.

Potassium Sulphate. *Potassii Sulphas.* K_2SO_4 .—This salt is not made directly, but is a by-product in the manufacture of nitric acid from potassium nitrate and sulphuric acid—e.g.: $2\text{KNO}_3 + \text{H}_2\text{SO}_4 = \text{K}_2\text{SO}_4 + 2\text{HNO}_3$. The HNO_3 distils over, leaving the potassium sulphate in the retort. Repeated crystallization purifies the salt. It is in colorless, six-sided prisms, permanent, odorless, sharp saline taste,

neutral reaction. It was formerly used as a vehicle or diluent in Dover's powder, but has been superseded by sugar of milk in the 1880 U. S. P.

The carbonate, chromate, cyanide, dichromate, ferricyanide, ferrocyanide, hydroxide, iodide, sulphate and sulphocyanate of potassium are official in form of test-solutions.

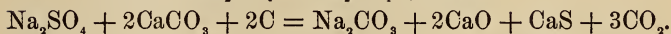
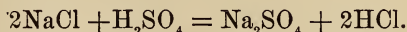
In the consideration of the potassium salts it has been thought best to follow the order of derivation, which necessitated the study of the several organic salts at the same time. The organic potassium salts are the acetate, bitartrate, tartrate, citrate and tartrate of sodium and potassium.

Sodium Salts.

THE sodium salts resemble the potassium salts very closely in their physical and chemical behavior and properties and in the methods of their preparation. There are no two other metals that resemble each other more closely in their combinations than do sodium and potassium; nearly all that has been said of the potassium salts is equally true of the sodium combinations. The latter are somewhat more soluble, so that the tartrate test and the platinic chloride test are not applicable. The only test for sodium consists in its imparting a bright-yellow color to a colorless flame. The general impurities and tests are the same as in the potassium compounds.

Sources.—The sodium chloride found in solution in brines and in the sea, and crystallized or in masses in mines, is the chief source of the sodium salts used in pharmacy. As with potassium, the carbonate is the salt from which

most others are made—this salt being made from the chloride by conversion. The processes employed for conversion are the *Leblanc* and the *Solvay* or *Ammonia* processes. The Leblanc process was employed very largely before the Solvay process was introduced, but is rapidly falling into disuse. It consists in treating the chloride with sulphuric acid and heating the resulting sodium sulphate with limestone, the combined action of which decomposes it to form sodium carbonate. The soluble sodium carbonate is afterward obtained by washing the residue with water; e.g.:



The Na_2CO_3 in crystallizing takes up 10 molecules of water and becomes the “soda” of commerce. In the Solvay process strong solutions of sodium chloride are treated simultaneously with ammonia, NH_3 , and carbonic acid gas, CO_2 , yielding *bicarbonate* of sodium: $\text{NaCl} + \text{NH}_3 + \text{CO}_2 + \text{H}_2\text{O} = \text{NaHCO}_3 + \text{NH}_4\text{Cl}$. The solutions are concentrated so that the bicarbonate precipitates and the ammonium chloride formed remains in solution. The dry ammonium chloride obtained by evaporating its solution is heated with lime, CaO , obtained in making the CO_2 for the process from CaCO_3 ($\text{CaCO}_3 + \text{heat} = \text{CaO} + \text{CO}_2$), and is in this way made to yield $2\text{NH}_4\text{Cl} + \text{CaO} = 2\text{NH}_3 + \text{CaCl}_2 + \text{H}_2\text{O}$. This process is very economical, it being only necessary to employ ammonia once. It furnishes a comparatively pure bicarbonate of sodium. Other natural salts of sodium besides the chloride are the nitrate or Chili saltpetre, found as an efflorescence on the soil in Chili, the borate or borax occurring crystallized and imbedded in the mud of lake

bottoms in California, Tuscany, etc., or as solution in borax-springs, and the sulphate in solution in springs. The nitrate and borate are never converted into other salts of sodium; the sulphate when it occurs naturally is usually the chief ingredient of the so-called *bitter waters*.

Cryolite, a double fluoride of sodium and aluminum, $6\text{NaF} \cdot \text{Al}_2\text{F}_6$, or $3\text{NaF} \cdot \text{AlF}_3$, found abundantly in Greenland, furnishes a considerable quantity of the sodium bicarbonate of the market to-day. The ore is melted with lime, yielding solution of aluminate of sodium and insoluble calcium fluoride: $6\text{NaF} \cdot \text{Al}_2\text{F}_6 + 6\text{CaO} = 6\text{CaF}_2 + 3\text{Na}_2\text{O} \cdot \text{Al}_2\text{O}_3$, or $\text{Na}_6\text{Al}_2\text{O}_6$. The solution of $3\text{Na}_2\text{O} \cdot \text{Al}_2\text{O}_3$ is treated with $6\text{CO}_2 + 6\text{H}_2\text{O}$ and yields $\text{Al}_2(\text{OH})_6 + 6\text{NaHCO}_3$. The aluminum hydroxide is insoluble, while the bicarbonate is removed with water. The bicarbonate thus produced is very pure. From the carbonate or bicarbonate of sodium all of the pharmacopœial sodium salts are made, excepting the borate and nitrate, which occur naturally, and the chlorate, sulphate and hypsulphite.

Sodium Carbonate. *Sodii Carbonas*. $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$.—Prepared, as above shown, by the Leblanc or Solvay processes. The latter process yields bicarbonate, which is converted into the carbonate by heating, to drive off part of the carbonic acid gas: $2\text{NaHCO}_3 + \text{heat} = \text{Na}_2\text{CO}_3 + \text{H}_2\text{O} + \text{CO}_2$. The salt crystallizes with 10 molecules of water, which it loses again by exposure to the air, it being efflorescent. Its chief use in pharmacy is in the preparation of other sodium salts, and in this application it is necessary to employ a salt containing the full amount of water of crystallization. The molecular weight of the salt when pure is 286, but if effloresced it is always stronger in

actual carbonate; and this must be taken into account when converting it into other salts.

Dried Sodium Carbonate. *Sodii Carbonas Exsiccatas*.—The water of crystallization is expelled by heat from the crystallized carbonate, so that 200 parts are reduced to 100 parts by weight.

The U. S. P. requires the presence of 73 per cent. anhydrous carbonate.

Sodium Bicarbonate, Commercial.—Made by Solvay process, or by passing a stream of carbonic acid gas into the carbonate: $\text{Na}_2\text{CO}_3 + \text{CO}_2 + \text{H}_2\text{O} = 2\text{NaHCO}_3$, e.g., not less than 95 per cent. of NaHCO_3 need be present, and not more than 5 per cent. impurities. The U. S. P. 1880 required this salt to be of 95 per cent. strength. It is no longer official.

Sodium Bicarbonate. *Sodii Bicarbonas*. NaHCO_3 .—By this title the U. S. P. designates a salt of 98.6 per cent. strength. It is made by removing the chloride, sulphate and carbonate from the commercial salt by washing the latter with a small quantity of water, which dissolve out the very soluble impurities before it dissolves the bicarbonate, which is less soluble.

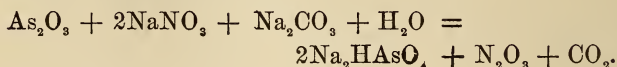
It is also made by the Solvay process and from Cryolite, as shown above. This purified salt is devoid of the bitter taste of the commercial, because of the absence of carbonate and sulphate.

Soda.—CAUSTIC SODA.—Sodium Hydroxide, NaOH . If sodium carbonate is heated with lime, CaO , the lime removes the CO_2 and leaves NaOH , sodium hydroxide, in solution, which, when evaporated to dryness and fused and poured into moulds, constitutes the stick form of caustic soda: $\text{Na}_2\text{CO}_3 + \text{CaO} + \text{H}_2\text{O} = \text{CaCO}_3 + 2\text{NaOH}$. When ex-

ceedingly pure caustic soda is required for analytical purposes, the metal sodium is oxidized, thus: $\text{Na}_2 + \text{O} = \text{Na}_2\text{O}$; $\text{Na}_2\text{O} + \text{H}_2\text{O} = 2\text{NaOH}$.

Solution of Soda. *Liquor Sodæ*.—Made either directly from the carbonate and lime, without evaporating the solution, or by dissolving the stick soda in water to the extent of 5 per cent. It is antacid, and resembles closely the Solution of Potassa.

Sodium Arsenate. *Sodii Arsenas*. $\text{Na}_2\text{HASO}_4 + 7\text{H}_2\text{O}$.—Prepared by fusing together sodium carbonate, sodium nitrate and arsenous acid, dissolving the fused mass in water and crystallizing:



This salt is sometimes preferred to arsenous acid; its medical properties are the same. Dose, 0.0027 to 0.008 Gm.

Sodium Bisulphite. *Sodii Bisulphis*. NaHSO_3 .—Sulphurous acid gas, SO_2 , is passed into solution of sodium carbonate to saturation, the solution evaporated and crystallized: $\text{Na}_2\text{CO}_3 + \text{H}_2\text{O} + 2\text{SO}_2 = 2\text{NaHSO}_3 + \text{CO}_2$. The sulphite, when exposed to the air, becomes sulphate by absorbing oxygen, or carbonate by taking up CO_2 ; it should therefore be kept in small bottles well filled. Its principal use is as an antiseptic.

Sodium Sulphite. *Sodii Sulphis*. $\text{Na}_2\text{SO}_3 + 7\text{H}_2\text{O}$.—The normal sodium sulphite is made by the same process as the bisulphite or acid sulphate, except that a greater quantity of SO_2 is employed: $\text{Na}_2\text{CO}_3 + \text{SO}_2 = \text{Na}_2\text{SO}_3 + \text{CO}_2$. This salt is a very useful antiferment and antiseptic, and is usually preferred to the bisulphite.

Sodium Bromide. *Sodii Bromidum*. NaBr .—Usually made

by acting on ferrous bromide, FeBr_2 , with sodium carbonate, filtering, evaporating the solution and crystallizing or granulating: $\text{FeBr}_2 + \text{Na}_2\text{CO}_3 = \text{FeCO}_3 + 2\text{NaBr}$. The ferrous bromide is prepared by acting on iron with bromine. Bromide of sodium is similar in medicinal properties to potassium bromide. Its dose is from 0.50 to 2.00 Gms.

Sodium Chloride. *Sodii Chloridum.* NaCl .—As above stated, this salt is the source of most of the sodium salts. It is obtained from sea-water and from salt-wells by evaporation, or from salt-mines, and purified by repeated crystallization. It is used most largely as a condiment, and is an essential constituent of our food. The presence of calcium or magnesium chloride causes the granulated salt to cake.

Sodium Hypophosphite. *Sodii Hypophosphis.* $\text{NaH}_2\text{PO}_2 + \text{H}_2\text{O}$.—Sodium carbonate is treated with calcium hypophosphite, when double decomposition ensues, forming insoluble calcium carbonate and soluble sodium hypophosphite, which is obtained by carefully evaporating the solution and granulating: $\text{Na}_2\text{CO}_3 + \text{Ca}(\text{H}_2\text{PO}_2)_2 = 2\text{NaH}_2\text{PO}_2 + \text{CaCO}_3$. The evaporation should be conducted at a very low heat to avoid explosion. Too much heat causes the disengagement of the inflammable gases, hydrogen and phosphoretted hydrogen. This salt is an ingredient in Syrup of Hypophosphites.

When this salt is heated or triturated with nitrates, chlorates or other oxidizing agents it explodes violently; hence it must be handled with extreme care.

Sodium Iodide. *Sodii Iodidum.* NaI .—This salt is prepared as the bromide, using ferrous iodide: $\text{FeI}_2 + \text{Na}_2\text{CO}_3 = \text{FeCO}_3 + 2\text{NaI}$. Also made by adding iodine to solution of soda: $6\text{NaOH} + 3\text{I}_2 = 5\text{NaI} + \text{NaIO}_3 + 3\text{H}_2\text{O}$. The iodate, NaIO_3 , is reduced to the iodide with charcoal,

viz.: $2\text{NaIO}_3 + 3\text{C}_2 = \text{Na}_2\text{I} + 6\text{CO}$. All that has been said about potassium iodide applies to this salt.

Sodium Phosphate. *Sodii Phosphas.* $\text{Na}_2\text{HPO}_4 + 12\text{H}_2\text{O}$.—This is an acid salt made by displacing part of the calcium in acid calcium phosphate with sodium from sodium carbonate, e.g.: $\text{CaH}_4(\text{PO}_4)_2 + \text{Na}_2\text{CO}_3 = \text{CaHPO}_4 + \text{Na}_2\text{HPO}_4 + \text{H}_2\text{O} + \text{CO}_2$. The CaHPO_4 is insoluble and is separated by filtration. This salt contains over 60% of water of crystallization. Used principally as a cathartic in doses of half an ounce, and in preparing the test-solution of sodium phosphate.

Sodium Pyrophosphate. *Sodii Pyrophosphas.* $\text{Na}_4\text{P}_2\text{O}_7 + 10\text{H}_2$.—By removing 1 molecule of water from 2 molecules of the acid phosphate, by heat, the pyrophosphate is obtained: $2\text{Na}_2\text{HPO}_4 + \text{heat} = \text{Na}_4\text{P}_2\text{O}_7 + \text{H}_2\text{O}$. It is employed in the preparation of ferric pyrophosphate.

Sodium Acetate. *Sodii Acetas.* $\text{NaC}_2\text{H}_3\text{O}_2 + 3\text{H}_2\text{O}$.—Obtained by adding acetic acid to solution of sodium carbonate until effervescence ceases, evaporating and crystallizing or granulating. $2\text{HC}_2\text{H}_3\text{O}_2 + \text{Na}_2\text{CO}_3 = 2\text{NaC}_2\text{H}_3\text{O}_2 + \text{H}_2\text{O} + \text{CO}_2$. As diuretic it is preferred by some to potassium acetate.

Sodium Benzoate. *Sodii Benzoas.* $\text{NaC}_7\text{H}_5\text{O}_2$.—Made as the above, using benzoic in place of acetic acid: $\text{Na}_2\text{CO}_3 + 2\text{HC}_7\text{H}_5\text{O}_2 = 2\text{NaC}_7\text{H}_5\text{O}_2 + \text{H}_2\text{O} + \text{CO}_2$. With ferric sulphate this salt gives a flesh-colored precipitate.

Sodium Salicylate. *Sodii Salicylas.* $\text{NaC}_7\text{H}_5\text{O}_3$.—Made as the above, employing salicylic in place of benzoic acid: $\text{Na}_2\text{CO}_3 + 2\text{HC}_7\text{H}_5\text{O}_3 = 2\text{NaC}_7\text{H}_5\text{O}_3 + \text{H}_2\text{O} + \text{CO}_2$. This salt gives with ferric compounds an intense violet color.

Sodium Sulphocarbolate. *Sodii Sulphocarbolas.* $\text{NaSO}_3\text{C}_6\text{H}_4(\text{OH}) + 2\text{H}_2\text{O}$. *Socium Paraphenolsulphonate.* Barium sul-

phocarbonate is treated with sodium carbonate, forming barium carbonate, which precipitates, and sodium sulphocarbonate, which remains in solution, from which it is obtained by crystallization. This salt is used as an antiferment and a disinfectant.

Sodium Borate. *Sodii Boras.* $\text{Na}_2\text{B}_4\text{O}_7 + 10\text{H}_2\text{O}$.—The source of borax has already been mentioned. It is sometimes made by fusing boric acid with dried sodium carbonate. A little added to alcohol containing sulphuric acid gives a green color to the flame of burning alcohol. Borax is mildly alkaline and diuretic. It is frequently used in solution to whiten ointments. See Boron.

Sodium Nitrate. *Sodii Nitras.* NaNO_3 .—Chili nitre, Chili saltpetre. Found native on the soil of Chili and Peru. Purified by repeated crystallization. Its chief use is in the manufacture of nitric acid.

Sodium Sulphate. *Sodii Sulphas.* $\text{Na}_2\text{SO}_4 + 10\text{H}_2\text{O}$.—Obtained as an intermediate product in the manufacture of sodium carbonate by the Leblanc process, and as a by-product in manufacture of nitric and hydrochloric acids. Glauber's salt, as it is frequently termed, rapidly effloresces in dry air. It is used largely in veterinary practice as a cathartic.

Sodium Chlorate. *Sodii Chloras.* NaClO_3 .—Made from potassium chlorate and sodium bitartrate: $\text{KClO}_3 + \text{NaHC}_4\text{H}_4\text{O}_6 = \text{NaClO}_3 + \text{KHC}_4\text{H}_4\text{O}_6$. The cream of tartar separates while the chlorate of sodium remains in solution and is obtained in crystals upon filtration and evaporation. It has an advantage over potassium chlorate in its greater solubility. *It should not be triturated with easily oxidizable bodies.* Accidents have occurred from its being rubbed with sulphur.

Sodium Hyposulphite. *Sodii Hyposulphis.* $\text{Na}_2\text{S}_2\text{O}_3 + 5\text{H}_2\text{O}$.

—The correct name for this salt is *sodium thiosulphate*. It is obtained by decomposing calcium thiosulphate with sodium sulphate: $\text{CaS}_2\text{O}_3 + \text{Na}_2\text{SO}_4 = \text{CaSO}_4 + \text{Na}_2\text{S}_2\text{O}_3$. The solution when filtered and evaporated yields crystals of sodium thiosulphate. A solution of this salt constitutes the photographer's "hypo," which dissolves the unaltered bromide or chloride in the film. In pharmacy it is chiefly employed in volumetric analysis.

Solution of Chlorinated Soda. *Liquor Sodæ Chloræ*.—La barraque's Solution. This bleaching solution is easily made by adding to a solution of carbonate of sodium some chlorinated lime, allowing the carbonate of calcium formed to subside, and decanting the clear solution: $2\text{Na}_2\text{CO}_3 + \text{CaCl}_2\text{Ca}(\text{ClO})_2 = 2\text{CaCO}_3 + 2\text{NaCl} + 2\text{NaOCl}$. It is used principally as a bleaching agent, and is usually substituted for javelle-water, which is made from potassium carbonate instead of sodium carbonate. It should contain at least 2.6% of available chlorine.

Solution of Sodium Arsenate. *Liquor Sodii Arsenatis*.—A solution containing one per cent of sodium arsenate.

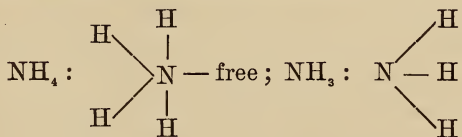
Solution of Sodium Silicate. *Liquor Sodii Silicatis*.—Made by fusing sand and sodium carbonate together, and dissolving the soluble portion of the mass thus obtained in water. It contains from 25 to 30 per cent. of silicate of sodium. Commercially it is known as *soluble glass*. It is employed in the preparation of some surgical dressings.

Troches of Sodium Bicarbonate. *Trochisci Sodii Bicarbonatis*.—These contain each three grains of sodium bicarbonate. Mild antacid.

Rhubarb and Soda Mixture. *Mistura Rhei et Sodæ*.—Contains about 3.5 per cent. of sodium bicarbonate.

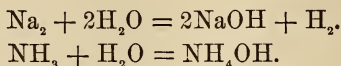
Ammonium Salts.

MUCH which the student has learned in the study of the potassium and sodium salts may be applied with profit in the study of the salts of ammonium. These salts are akin in many respects to those of potassium and sodium, especially in absence of color, solubility, physical and chemical properties and medicinal uses. The base ammonium is, however, of a compound nature, consisting of nitrogen and hydrogen, NH_4 , while the salts of sodium and potassium have single elements as their bases. This, with the fact that ammonium salts are all volatile, marks the chief points of distinction between them. The student should be careful not to mistake ammonium, NH_4 , for ammonia, NH_3 . The former is a hypothetical base, which has not yet been obtained by itself, or isolated from any of its compounds, while the latter is a gas easily generated from any ammonium salt by heating with an alkali hydroxide. We assume that ammonium has one free bond, and ammonia none, and that the nitrogen in the former has five bonds and in the latter only three:



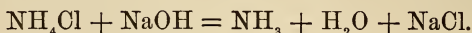
NH_4 is an unsaturated group of atoms having one free bond (the nitrogen having 5 bonds, of which 4 are satisfied with 4 hydrogen atoms), and does exist by itself, thus differing from NH_3 , which is completely saturated, and a stable though extremely volatile body. The latter, NH_3 , differs from sodium or potassium in uniting or com-

binning with *all* of the elements of another body. The following equations illustrate this difference:



The sodium, it will be observed, does not unite with all of the hydrogen in water; it sets one atom from each molecule free. Ammonia, on the contrary, enters into complete combination, the number of bonds of the nitrogen increasing from 3 to 5.

Tests for Ammonium Salts.—1. The presence of *any* ammonium salt may be easily detected by heating with caustic soda or potassa, or with any alkali hydroxide. The gas ammonia is given off, which may be recognized by its odor, or by its producing white fumes when brought in contact with the vapors of hydrochloric acid. For illustration:



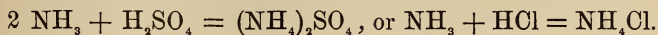
3. Solutions of ammonium salts give yellow precipitates with solution of platinic chloride containing a little HCl.

4. Heat volatilizes ammonium salts.

5. Nessler's reagent gives red color with even minute quantities of ammonia. (See page 45.)

Sources.—In one of the processes for the manufacture of illuminating gas, coal is subjected to destructive distillation; the gas thus generated contains among other products considerable quantities of ammonia. This is removed from the illuminating gas in the process of washing, the water dissolving out the NH_3 and retaining it in solution. From this solution, or *ammoniacal gas liquor*, as it is called, the ammonia is removed by treatment with either sulphuric

or hydrochloric acid, evaporating and purifying by sublimation the sulphate or chloride thus formed:



If sulphuric acid has been employed, the sulphate formed is converted into the chloride by subliming with sodium chloride: $2\text{NaCl} + (\text{NH}_4)_2\text{SO}_4 = \text{Na}_2\text{SO}_4 + 2\text{NH}_4\text{Cl}$.

Ammonium Chloride. *Ammonii Chloridum*. NH_4Cl .—This is the starting-point in the manufacture of the ammonium salts; from it all others are made, either directly or indirectly, through the sulphate or carbonate or hydroxide. The impurities found in this salt may be looked for in all the others. They will be mentioned here only, and are sulphate, chloride, metals, iron especially, from the subliming apparatus, and empyreumatic substances. The latter are present if a solution of a salt discharges the color of potassium permanganate in presence of sulphuric acid.

The crude ammonium chloride, or *sal-ammoniac*, as it is called, usually occurs in tough, fibrous masses, or, when purified, in granular form. The former is used largely in batteries, and the latter for medicinal purposes.

Ammonia-Water. *Aqua Ammoniacæ*. NH_4OH .—This preparation is a solution of ammonia, NH_3 , in water to the extent of 10 per cent. by weight. It is made by heating pieces of lime, water and ammonium chloride in a retort and conducting the ammonia gas into water: $\text{H}_2\text{O} + \text{CaO} = \text{Ca}(\text{OH})_2$. $\text{Ca}(\text{OH})_2 + 2\text{NH}_4\text{Cl} = \text{CaCl}_2 + 2\text{H}_2\text{O} + 2\text{NH}_3$. $\text{NH}_3 + \text{H}_2\text{O} = \text{NH}_4\text{OH}$. It is more conveniently made by diluting the

Stronger Water of Ammonia, *Aqua Ammoniacæ Fortior*, which is made as the above, the gas being passed into water until the latter is saturated, or until it contains 28 per cent. by

weight of the gas. The two preparations differ only in strength. They both have an excessively pungent odor, acrid and alkaline taste, and strongly alkaline reaction. They differ chemically from the potass. and sod. solutions in being volatile and in emulsifying fats instead of saponifying them. The specific gravity of the stronger water is 0.90; that of the weaker, 0.96. Both should be kept in well-stoppered bottles in cool places.

Ammonia Liniment. *Linimentum Ammoniae, Volatile Liniment.*—Cotton-seed oil is emulsified with water of ammonia and a little alcohol added. The liniment should always be well shaken before being applied, and should be made when wanted. The alcohol aids in binding together the oil and the solution.

Spirit of Ammonia. *Spiritus Ammoniae.*—Spirit of ammonia is a solution of 10 per cent. of NH_3 in alcohol, and is obtained by heating the stronger water of ammonia and conducting the gas thus volatilized into alcohol. Ammonia-water is often, but erroneously, called “spirit of ammonia.”

Aromatic Spirit of Ammonia. *Spiritus Ammoniae Aromaticus.*—This is a solution of ammonium carbonate and ammonia-water in alcohol and water, and flavored with the oils of nutmeg, lemon and lavender-flowers. It is intended for internal use, and is a very valuable stimulant and antacid. When first made it is colorless, but it gradually becomes darker, owing to the action of the ammonia upon the oils and upon the alcohol. Precipitation often takes place, which is due to the ammonium carbonate employed. The official salt contains some bicarbonate, which the added ammonia-water changes to the carbonate; but if the official salt has been exposed, or if it has effloresced, it contains more bicarbonate than the ammonia-water can convert into car-

bonate, and it is this excess of ammonium bicarbonate which forms the precipitate which sometimes deposits in the preparation, it being insoluble in the mixture of alcohol and water.

Ammonium Carbonate. *Ammonii Carbonas.* $\text{NH}_4\text{HCO}_3 \cdot \text{NH}_4 \cdot \text{NH}_2\text{CO}_2$.—This salt is not a true carbonate, but consists of a mixture of one molecule of *bicarbonate* and one of *carbamate*. It is made from calcium carbonate and ammonium chloride; in this process double decomposition takes place according to the following reaction: $2\text{CaCO}_3 + 4\text{NH}_4\text{Cl} = 2\text{CaCl}_2 + \text{H}_2\text{O} + \text{NH}_3 + (\text{NH}_4\text{HCO}_3 \cdot \text{NH}_4 \cdot \text{NH}_2\text{CO}_2)$. If one molecule of water is added to the carbamate a neutral carbonate is produced, thus: $\text{NH}_4 \cdot \text{NH}_2\text{CO}_2 + \text{H}_2\text{O} = (\text{NH}_4)_2\text{CO}_3$. This happens when the pharmacopœial salt is dissolved in water, the resulting solution containing the acid and neutral carbonates. By exposure to air or long keeping the salt loses ammonia and carbonic acid gas, the proportion of bicarbonate thereby increasing. Ammonia-water converts the bicarbonate into the carbonate.

Ammonium carbonate is usually in form of translucent masses, of very pungent and ammoniacal odor, alkaline, and partly soluble in alcohol, the bicarbonate remaining undissolved. The chief and objectionable impurity in this salt is empyreumatic matter, which communicates to the salt a smoky odor, and whose presence is an indication of incomplete purification.

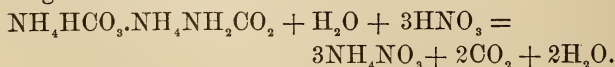
Solution of Ammonium Acetate. *Liquor Ammonii Acetas.* $\text{NH}_4\text{C}_2\text{H}_3\text{O}_2$.—Ammonium carbonate is employed in making this solution, which is often called Spirit of Mindererus. It is neutralized by dilute acetic acid: $\text{NH}_4\text{HCO}_3 \cdot \text{NH}_4 \cdot \text{NH}_2\text{CO}_2 + \text{H}_2\text{O} + 3\text{HC}_2\text{H}_3\text{O}_2 = 3\text{NH}_4\text{C}_2\text{H}_3\text{O}_2 + 2\text{CO}_2 + 2\text{H}_2\text{O}$.

This preparation should be freshly prepared when needed, and should be very slightly acid.

Ammonium Benzoate. *Ammonii Benzoas.* $\text{NH}_4\text{C}_7\text{H}_5\text{O}_2$.—Conveniently made by dissolving benzoic acid in ammonia-water, or by adding benzoic acid to solution of ammonium carbonate. $\text{NH}_4\text{OH} + \text{HC}_7\text{H}_5\text{O}_2 = \text{NH}_4\text{C}_7\text{H}_5\text{O}_2 + \text{H}_2\text{O}$. The benzoate should be crystallized from a slightly alkaline solution to prevent the formation of the acid benzoate. The salt emits ammonia and benzoic acid vapors when strongly heated, and with solution of sulphate of iron containing a little nitric acid it produces a flesh-colored precipitate. By treatment with an acid the benzoic acid separates.

Ammonium Bromide. *Ammonii Bromidum.* NH_4Br .—This salt may be made by dissolving bromine in ammonia-water or by a double decomposition between ammonium sulphate and potassium bromide: $(\text{NH}_4)_2\text{SO}_4 + 2\text{KBr} = 2\text{NH}_4\text{Br} + \text{K}_2\text{SO}_4$. The addition of alcohol to the solution causes the potassium sulphate to precipitate. This salt is sometimes made from ferrous bromide and ammonium carbonate. See p. 185. Ammonium bromide loses some of its iodine upon exposure and becomes yellow. The Pharmacopœia allows the presence of one per cent. chloride.

Ammonium Nitrate. *Ammonii Nitras.* NH_4NO_3 .—If ammonium carbonate is treated with diluted nitric acid until effervescence ceases and the solution evaporated, crystals of nitrate of ammonium will be formed. Sometimes the evaporation is carried to dryness, to obtain the salt in masses or in granular form:



The chief use of this salt is to prepare the nitrous oxide

used as an anesthetic in dentistry. The laughing-gas is obtained by simply heating the nitrate: $\text{NH}_4\text{NO}_3 + \text{heat} = \text{N}_2\text{O} + 2\text{H}_2\text{O}$.

Ammonium Valerianate. *Ammonii Valerianas.* $\text{NH}_4\text{C}_5\text{H}_9\text{O}_2$.—Double decomposition between valeric or valerianic acid and ammonia-water or ammonium carbonate produces ammonium valerianate. Among the several methods employed in making this salt the best is to pass ammonia gas into solution of valeric acid: $\text{HC}_5\text{H}_9\text{O}_2 + \text{NH}_3 = \text{NH}_4\text{C}_5\text{H}_9\text{O}_2$. It occurs in white flakes, is deliquescent, has strong odor of valeric acid and is usually of acid reaction, though it should be neutral. It is a nervine and is largely used to prepare the popular elixir of ammonium valerianate. The commercial salt usually contains an excess of acid. It is very soluble in alcohol.

Ammonium Iodide. *Ammonii Iodidum.* NH_4I .—Obtained by double decomposition between potassium iodide and ammonium sulphate: $(\text{NH}_4)_2\text{SO}_4 + 2\text{KI} = 2\text{NH}_4\text{I} + \text{K}_2\text{SO}_4$. The potassium sulphate is made to crystallize first from the solution by the addition of alcohol and reduction of temperature. The salt is not permanent in the air, soon becoming yellow from the liberation of iodine. It should be white and odorless. It is soluble in alcohol, and is used in much the same indications as potassium iodide.

Lithium Salts.

LITHIUM is the last of the alkalies left for study. The metal is peculiar in that it is the lightest of all known metals, being a little more than half as heavy as water; specific gravity, 0.589. Its salts resemble very closely those of potassium and sodium; they are all white or colorless, soluble in water, excepting the carbonate, which is only

slightly soluble. The benzoate and carbonate are permanent in air; the others are very deliquescent. Lithium is widely distributed in nature, but always occurs in very small quantities. The impurities likely to be found are principally salts of other alkalies and of alkaline earths. The sources of the lithium salts are at present the minerals *spodumene* and *lepidolite*, from which, by a circumstantial process, the carbonate is prepared, this in turn yielding all the other salts.

Lithium and its salts are easily detected by the flame, to which they give a vivid red color. Sodium phosphate or ammonium carbonate gives a white precipitate in concentrated solutions of lithium salts of neutral or alkaline reaction. The precipitate is dissolved by acids or by ammonia-water. Lithium is a monad; atomic weight, 7.01.

Lithium Carbonate. *Lithii Carbonas.* Li_2CO_3 .—This is the salt obtained from the natural combinations of lithium, by first converting the latter into the sulphate, and precipitating the lithium carbonate by ammonia carbonate. It dissolves in acids with effervescence, producing corresponding salts; is soluble in 80 parts of water and insoluble in alcohol. It is frequently administered in cases of gouty affections.

Lithium Benzoate. *Lithii Benzoas.* $\text{LiC}_7\text{H}_5\text{O}_2$.—Is prepared by heating together lithium carbonate and benzoic acid until effervescence ceases, filtering, evaporating and crystallizing. Sometimes the evaporation is carried on to dryness: $\text{Li}_2\text{CO}_3 + 2\text{HC}_7\text{H}_5\text{O}_2 = 2\text{LiC}_7\text{H}_5\text{O}_2 + \text{CO}_2 + \text{H}_2\text{O}$. The benzoate is soluble in water and alcohol, is of slightly acid reaction, and has an odor of benzoic acid. Its principal use is in the treatment of gout, although it is also used extensively as a means of dissolving urinary calculi.

Lithium Bromide. *Lithii Bromidum.* LiBr .—The bromide

results when ferrous bromide is decomposed with lithium carbonate, lithium bromide remaining in solution while the ferrous carbonate precipitates—e.g., $\text{FeBr}_2 + \text{Li}_2\text{CO}_3 = \text{FeCO}_3 + 2\text{LiBr}$. This salt is the most deliquescent of the lithium salts, is very soluble in alcohol, and is a reliable hypnotic administered in doses of 0.50 to 2.00 gms.

Lithium Citrate. *Lithii Citras.* $\text{Li}_3\text{C}_6\text{H}_5\text{O}_7$.—If to a warm solution of citric acid lithium carbonate is added as long as effervescence lasts, a solution of lithium citrate will be formed from which the salt is obtained by evaporating to dryness: $2\text{H}_3\text{C}_6\text{H}_5\text{O}_7 + 3\text{Li}_2\text{CO}_3 = 2\text{Li}_3\text{C}_6\text{H}_5\text{O}_7 + 3\text{CO}_2 + 3\text{H}_2\text{O}$. The citrate is used as the carbonate is and for same medicinal purpose.

Lithium Salicylate. *Lithii Salicylas.* $\text{LiC}_7\text{H}_5\text{O}_3$.—If instead of using citric acid, salicylic acid is employed, lithium salicylate results: $\text{Li}_2\text{CO}_3 + 2\text{HC}_7\text{H}_5\text{O}_3 = 2\text{LiC}_7\text{H}_5\text{O}_3 + \text{H}_2\text{O} + \text{CO}_2$. This salt is often used in place of sodium salicylate for rheumatism.

The Metals of the Alkaline Earths.

THE three metals constituting this group, calcium, barium and strontium, have some similarity to the alkalies, especially in their action upon vegetable colors when in form of oxide or hydroxide, whence their name, *alkaline* earths. They have also a similar taste and caustic action. The three metals are each dyad; they form only one chloride each, and their carbonates and sulphates are practically insoluble.

Calcium Salts.

THIS metal occurs very abundantly and widely diffused in nature, but never in the free state. The metal is not

unlike gold in some respects physically—it has a light-yellow color, is very ductile, easily cut or filed, and capable of being hammered into very thin sheets. It differs from gold in being much lighter in weight, in tarnishing in the air, and in being quickly acted upon by water and dilute acids.

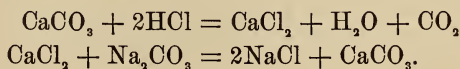
Sources.—The chief sources of calcium and its salts are the carbonate, which occurs in a great variety of forms: limestone, marble, chalk, Iceland spar, calcite, stalagmites, shells of oysters and of other cretaceous animals; the phosphate in the mineral apatite and in bones of vertebrates; and the sulphate as gypsum (native calcium sulphate). There are many other forms of occurrence, but the above are the more important ones.

Prepared Chalk. *Creta Præparata*. CaCO_3 .—The greater number of the pharmacopœial salts of calcium are made from the carbonate, which occurs in the numerous forms mentioned above. The native friable carbonate, known as chalk, is purified simply by a process of elutriation, which consists of triturating the chalk to a very fine powder in a slow stream of water, which flows into a series of tanks, each of which is somewhat lower than the one next to it, so that each becomes filled by the overflow of the next higher one. The coarser and heavier particles are deposited in the first tank, the lighter ones passing with the overflow into the next, where more of the heavier particles settle; the last tank usually has deposited in it very fine particles of chalk, which constitute the prepared chalk of the U. S. P. While still in the moist condition it is sometimes shaped into small cones, when it forms the “drop chalk” of the stores. This preparation is used in preference to any other kind of carbonate in medicine. It is antacid, and enters into

the *Compound Chalk Powder*, which contains in addition some acacia and powdered sugar, and which, when suspended in cinnamon-water, constitutes the popular *Chalk Mixture* of the U. S. P.

Prepared chalk may contain traces of barium, strontium, iron and magnesium as impurities. It is readily soluble in hydrochloric acid.

Precipitated Calcium Carbonate. *Calcii Carbonas Præcipitatus.* CaCO_3 .—The prepared chalk is not very pure chemically, so that the U. S. P. adopted another form of the carbonate, made by precipitating calcium *chloride* with sodium carbonate. The chloride is made from the carbonate obtained in nature by treatment with hydrochloric acid; the process therefore begins with a carbonate and ends with a carbonate:



This process yields a very finely-divided powder, usually very pure. It is less valuable than the prepared chalk for internal administration, but is used very largely in tooth-powders.

Troches of Chalk.—Chalk troches contain each 4 grains of prepared chalk, besides powdered gum arabic, sugar and nutmeg, made into a mass with water.

Calcium Chloride. *Calcii Chloridum.* CaCl_2 .—Made, as shown above, from chalk and hydrochloric acid. It is very deliquescent, and is chiefly used in chemical operations for removing watery vapors from gases. It is also largely used for keeping the air dry in glass-cased balances or electric-machines. It should be kept in bottles securely corked.

Calcium Bromide. *Calcii Bromidum.* CaBr_2 .—Of several

methods for preparing this salt, the simplest consists in adding precipitated chalk in excess to hydrobromic acid, filtering, evaporating to dryness or granulating: $\text{CaCO}_3 + 2\text{HBr} = \text{CaBr}_2 + \text{H}_2\text{O} + \text{CO}_2$. This is a very deliquescent salt, and needs to be kept in well-stoppered bottles. Its value depends on the bromine which it contains; it is a hypnotic and is given in doses of 0.50 to 2.00 gms.

Lime. *Ca/x.* CaO .—This is oxide of calcium, CaO , and is made by expelling the CO_2 from the carbonate by means of heat: $\text{CaCO}_3 + \text{heat} = \text{CaO} + \text{CO}_2$. Its chief use is in building; in pharmacy it furnishes lime-water, chlorinated lime, sulphurated lime and syrup of lime. It is often used for dehydrating purposes, as the chloride is; when it comes in contact with moisture or water it unites with it, producing the hydroxide: $\text{CaO} + \text{H}_2\text{O} = \text{Ca(OH)}_2$. Externally it is a caustic and an escharotic, and enters into preparations for removing hair from the skin.

Lime-Water. *Liquor Calcis.* Ca(OH)_2 .—Lime is slightly soluble in water, and when dissolved in it to saturation furnishes lime-water, which contains about 0.17 per cent of calcium hydroxide, Ca(OH)_2 . It is peculiar with many calcium salts that they are more soluble in cold water than in hot, and in making lime-water cold water should invariably be employed, since it is desired to get the largest amount of lime in solution. It is given to children to check nausea.

Syrup of Lime. *Syrupus Calcis.*—Lime is more soluble in a solution of sugar than in water alone, and syrup of lime is a solution of lime made by boiling lime and sugar in water. It contains about 6.5 per cent of lime, which is a greater percentage than is in the *Liq. Calcis*, and hence it is more alkaline.

Lime Liniment. *Linimentum Calcis*.—Equal parts of lime-water and cottonseed oil well shaken together constitute the lime liniment which is so largely used as application to burns, scalds, etc. The physicians prescribe it at times under the name “carron oil.”

Sulphurated Lime. *Calx Sulphurata*. *Crude Calcium Sulphide*, CaS , and CaSO_4 .—About equal parts of lime and precipitated sulphur are fused together in a crucible for an hour or longer, and the resulting product preserved in small, well-stoppered bottles. A chemical combination takes place producing varying quantities of calcium monosulphide and sulphate, but at least 60 per cent of the former should be present. This preparation when made into paste with water is an excellent depilatory. In small proportions it is sometimes used in ointments for skin-diseases, and at times internally for a similar purpose.

The U. S. P. of 1890 directs this preparation to be made by reducing the calcium sulphate to sulphide with charcoal and starch. Some of the sulphate remains unchanged.

Chlorinated Lime. *Calx Chlorata*. $3\text{Ca}(\text{OH})_2$.—Chemically this is a mixture of calcium chloride and hypochlorite: $2\text{CaO} + 2\text{Cl}_2 + \text{H}_2\text{O} = (\text{CaCl}_2 + \text{Ca}(\text{ClO})_2) + \text{H}_2\text{O}$. It is a grayish-white powder, should not be more than slightly damp, has an odor of chlorine, and should be kept in securely-closed vessels. It is employed very extensively in the arts for bleaching and in the household for disinfecting. Its value always depends upon the amount of available chlorine which it contains, and which should never be less than 35 per cent. See page 78.

Labarraque's Solution has received mention under chlorine; it is prepared by acting on chlorinated lime with soda carbonate. See page 80.

Calcium Hypophosphite. *Calcii Hypophosphis.* $\text{Ca}(\text{H}_2\text{PO}_2)_2$.—The making of this salt involves some danger, owing to the employment of phosphorus and because of the generation of the spontaneously-inflammable phosphoretted hydrogen. Calcium hydroxide (slaked lime) suspended in water is boiled with phosphorus and the gases formed conducted by means of a hood or draught into the chimney, or into the air, as quickly as they are generated: $3\text{Ca}(\text{OH})_2 + 8\text{P} + 6\text{H}_2\text{O} = 2\text{PH}_3$, phosphoretted hydrogen, $+ 3\text{Ca}(\text{H}_2\text{PO}_2)_2$, calcium hypophosphite. A little of the insoluble phosphate, $\text{Ca}_3(\text{PO}_4)_2$, is formed, which is removed by filtration. The solution is evaporated and the salt granulated or allowed to crystallize. In evaporating extreme care should be taken not to permit the temperature to rise above 180°F . for fear of explosion. The salt should never be kept in warm places, as it is very liable to explode and disengage the powerfully-poisonous PH_3 . This is a salt which the pharmacist is wise to buy from a reliable house. It enters into the *Syrup of Hypophosphites*, and is usually employed to make the other hypophosphites of the U. S. P. and of the market.

Precipitated Calcium Phosphate. *Calcii Phosphas Præcipitatus.* $\text{Ca}_3(\text{PO}_4)_2$.—Calcined bones are dissolved in hydrochloric acid and the solution precipitated with ammonia-water, the precipitate thoroughly washed or strained and dried. Bones contain the phosphate of calcium, and in order to obtain this in a finely-divided state it is caused to enter solution, from which it is precipitated as a fine powder. It is a permanent powder, insoluble, and used for distributing the oils in making the aromatic waters. It is better adapted for this purpose than magnesium carbonate, because of its greater insolubility. Its principal use was in preparing the

Syrup of Lactophosphate of Calcium, U. S. P. 1880, which was made by dissolving the freshly-prepared phosphate in lactic acid and making into syrup with sugar and flavoring with orange-flower water. The 1890 U. S. P. directs the precipitated *carbonate* in place of the phosphate in making the syrup of lactophosphate of calcium.

Strontium Salts.

THE strontium salts have lately been introduced into therapeutics, and they seem designed to occupy a very prominent place in the modern *materia medica*. Their use in medicine was so long deferred because of the belief that they were poisonous. The French physiologist, Dr. Laborde, however, lately proved that the toxic properties of salts of strontium were due to the presence of the poisonous barium, and that when the latter is removed the pure strontium salts are therapeutic agents of such value as to likely replace, it is predicted by some, the alkali and alkaline-earth salts in medicine. The first to prepare chemically pure strontium salts was the French chemist Paraf-Javal. All the natural compounds of strontium: strontianite, a carbonate; celestine, a sulphate, and others, contain barium, which is removed with difficulty in the process of purification. Every pharmacist should therefore examine his strontium salts for barium. This is easily accomplished by adding a few drops of dichromate of potassium test-solution (a ten per cent solution) to a solution of one part salt in 20 of distilled water—a turbidity would indicate the presence of barium and the need to reject the salt.

Properties and Tests for Identity and Purity.—The

Pharmacopœia recognizes three strontium salts, the bromide, the iodide and the lactate. They are all hydrous, the lactate containing three molecules of water and the bromide and iodide six each. The application of heat at first melts the salts, then fuses them causing the loss of their water of crystallization; a higher heat decomposes them. They impart an intense red color to a non-luminous flame. They are easily soluble in water; the lactate is permanent in the air, while the bromide and iodide are very deliquescent and prone to decomposition, for which reason they should be kept in well-stoppered bottles.

Each of the salts in solution is precipitated as sulphate with the solution of any soluble sulphate. Calcium sulphate is very sparingly soluble in water, but yet its solution will precipitate strontium, showing that the strontium sulphate is less soluble than the calcium sulphate.

Potassium chromate solution forms a yellow precipitate of strontium chromate in solutions of strontium salts. Soluble carbonates also precipitate soluble strontium salts.

A 1-in-20 aqueous solution should not be affected by hydrogen sulphide, either before or after acidulation with a drop of hydrochloric acid—this would indicate the absence of arsenic, lead, copper. Ammonium sulphide test-solution (H_2S passed into ammonia-water to saturation) should not alter the aqueous solution, which would indicate the absence of iron, aluminum and other metals.

Strontium Bromide. *Strontii Bromidum.* $\text{SrBr}_2 + 6\text{H}_2\text{O}$.—This salt may be obtained by treating hydrobromic acid with strontium carbonate, filtering, evaporating and crystallizing: $\text{SrCO}_3 + 2\text{HBr} = \text{SrBr}_2 + \text{H}_2\text{O} + \text{CO}_2$. The carbonate should be pure. The bromide also results if strontium is burned in bromine vapor. The salt is odor-

less, has a bitter alkaline taste, and occurs in forms of colorless, transparent, six-sided crystals. Properties are given above.

Medicinally the bromide is recommended as a substitute for the bromides of other bases in the treatment of epilepsy and allied conditions.

Strontium Iodide. *Strontii Iodidum.* $\text{SrI}_2 + 6\text{H}_2\text{O}$.—If a solution of hydriodic acid is saturated with strontium hydroxide, the strontium iodide results, which may then be obtained by evaporation: $2\text{HI} + \text{Sr}(\text{OH})_2 = \text{SrI}_2 + 2\text{H}_2\text{O}$. The iodide is in transparent, colorless plates, odorless, bitterish taste, deliquescent. It turns yellow if exposed to air and light, due to the liberation of iodine, and hence should be kept well preserved in small dark, amber-colored glass-stoppered bottles. See properties above.

Medicinally the iodide is claimed to have the alterative action of iodides, but without depressing the general nutrition, or irritating the intestinal canal.

Strontium Lactate. *Strontii Lactas.* $\text{Sr}(\text{C}_3\text{H}_5\text{O}_3)_2 + 3\text{H}_2\text{O}$.—Strictly pure strontium carbonate is dissolved by means of gentle heat in absolute lactic acid diluted with water: $\text{SrCO}_3 + 2\text{HC}_3\text{H}_5\text{O}_3 = \text{Sr}(\text{C}_3\text{H}_5\text{O}_3)_2 + \text{H}_2\text{O} + \text{CO}_2$. The salt is obtained from the solution by evaporating and granulating. The salt occurs in form of a white granular powder or crystalline nodules, odorless, permanent in the air. For other properties and tests see above.

The lactate is claimed to be a successful remedy in the treatment of nephritis and chronic albuminuria. Dose, 0.65 to 1.95 Gm.

Barium.

THE salts of this metal are not of sufficient interest to the pharmacist to merit more than mention. The chloride of barium is a very useful reagent in solution for sulphates, the sulphate of barium being wholly insoluble: $\text{BaCl}_2 + \text{Na}_2\text{SO}_4 = \text{BaSO}_4 + 2\text{NaCl}$.

The dioxide, BaO_2 , is employed in the preparation of peroxide of hydrogen, and the nitrate and hydroxide dissolved in water are used as test solutions.

Magnesium Salts.

MAGNESIUM occurs in a great variety of combinations and is widely distributed throughout nature, but pharmacy derives its supply at present chiefly from the silicious hydrate found in the serpentine rocks of New Jersey, Pennsylvania and Ohio. The principal source formerly was the sulphate obtained from the Epsom Springs, in England.

The word "magnesia" is derived from the name of a town in Asia Minor in the neighborhood of which two minerals were found in abundance, one of a white color named *magnesia alba*, and the other of a blackish color termed *magnesia nigra*. The former was that which we now know as magnesium carbonate, while the latter is the manganese dioxide, called manganese to distinguish it from magnesium.

Tests for Magnesium Salts.—Sodium phosphate solution produces, in presence of ammonia-water, a white crystalline precipitate of NH_4MgPO_4 , ammonio-magnesium phosphate.

The caustic alkalies produce white gelatinous magma-like precipitates. The presence of ammonium chloride prevents the formation of these precipitates.

Any soluble carbonate gives with soluble magnesium salts a copious white precipitate.

Magnesium Sulphate. *Magnesii Sulphas.* $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$.—The sulphate is made from the mineral *serpentine* by digesting the latter with sulphuric acid, dissolving out the sulphate produced and crystallizing. Some years ago it was almost exclusively obtained by evaporating the water from Epsom Spring, in England, which contained a large proportion of the salt in solution. All other salts are made from the sulphate. It is in small colorless prisms, or acicular needles, efflorescent in dry air, very soluble in water, insoluble in alcohol. The *impurities* likely to be present, and consequently likely to be present in all of the salts made from it, are chlorides, metals, alkaline earths and alkalis; the U. S. P. permits the presence of 1 per cent of the sulphates of the alkalis. Epsom salt is purgative in 1 or 2-ounce doses. Oxalic acid is sometimes mistaken for it, although there is a great dissimilarity between the crystals.

Magnesium Carbonate. *Magnesii Carbonas.* (Approximately MgCO_3)₄. $\text{Mg}(\text{OH})_2 \cdot 5\text{H}_2\text{O}$.—The composition of magnesium carbonate varies somewhat. It always contains hydroxides, and the salt recognized by the U. S. P. probably contains one molecule of the hydroxide and four of the carbonate, as the formula indicates. There are two carbonates, a *light* and a *heavy* one, both made by the same process, the difference being merely in the employment of dilute or concentrated solutions of sodium carbonate and magnesium sulphate: $5\text{MgSO}_4 + 5\text{Na}_2\text{CO}_3 + \text{H}_2\text{O} = (\text{MgCO}_3)_4 \cdot \text{Mg}(\text{OH})_2 + 5\text{Na}_2\text{SO}_4 + \text{CO}_2$. The carbonate precipitates and, when thoroughly washed on a muslin or calico strainer, is dried. Concentrated solutions produce the heavy carbonate,

which is that recognized by the Pharmacopœia, and which occurs as a light white powder or in friable blocks or masses, insoluble, odorless, tasteless. It is employed in making the heavy magnesia, solution of magnesium citrate, mixture of magnesia and asafoetida of the 1880 U. S. P., etc. It is sometimes employed to diffuse the oils in making the medicated waters. Medicinally it is an antacid.

Magnesia. *Magnesia.* MgO .—This is the *light magnesia* or *calcined magnesia*, made from *light* carbonate by expelling the CO_2 : $\text{MgCO}_3 + \text{heat} = \text{MgO} + \text{CO}_2$. It is a very light, white, insoluble powder, very bulky, odorless, earthy taste, and alkaline when moistened with water. Magnesia is laxative, and is usually combined with powdered rhubarb.

Heavy Magnesia. *Magnesia Ponderosa.* MgO .—This corresponds in composition and all properties to the above, differing from it only in being heavier. It is made by expelling the CO_2 from the heavy carbonate.

Solution of Magnesium Citrate. *Liquor Magnesii Citratis.*—This very popular refrigerant laxative is a freshly prepared solution of magnesium citrate flavored with lemon and sugar, and made refrigerant and sparkling by confining the CO_2 disengaged from potassium bicarbonate: $3\text{MgCO}_3 + 2\text{H}_3\text{C}_6\text{H}_5\text{O}_7 = \text{Mg}_3(\text{C}_6\text{H}_5\text{O}_7)_2 + 3\text{H}_2\text{O} + 3\text{CO}_2$. Sometimes the calcined magnesia is employed in place of the carbonate; the disengagement of the CO_2 is thereby obviated. The solution is kept in strong 12-ounce bottles, with the cork tied tightly over the neck of the bottle to prevent its being forced out by the CO_2 from the potassium bicarbonate. The latter coming in contact with the citric acid, which is always in excess, loses its CO_2 by double composition accord-

ing to this reaction: $3\text{KHCO}_3 + \text{H}_3\text{C}_6\text{H}_5\text{O}_7 = \text{K}_3\text{C}_6\text{H}_5\text{O}_7$ (potassium citrate) $+ 3\text{H}_2\text{O} + 3\text{CO}_2$.

Effervescent Magnesium Citrate. *Magnesii Citras Effervescens*.—This effervescent salt is made by mixing intimately magnesium carbonate, sodium bicarbonate, powdered citric acid and sugar, and rubbing to a paste with alcohol. This paste is then passed through a coarse sieve and dried. When brought into contact with water, solution of the acid and bicarbonate takes place, and they react upon each other and upon the carbonate and produce effervescence through the liberation of CO_2 . It is essential to keep this preparation in well-closed bottles.

Magnesium Sulphite. $\text{MgSO}_3 \cdot 6\text{H}_2\text{O}$.—Prepared by passing pure sulphurous acid gas into magnesia suspended in water: $\text{MgO} + \text{SO}_2 = \text{MgSO}_3$. The sulphite is less soluble than the sulphate. It is employed very much as the sodium and potassium sulphites are. It is a white crystalline powder which oxidizes by exposure to air. It is no longer official.

Zinc Salts.

THE metal zinc is official in form of irregular granulated pieces or in thin sheets, or moulded into thin pencils, or in a state of fine powder. It is obtained from *calamine*, the chief source of the zinc salts, by roasting with charcoal and collecting the zinc vapors in water. It is a dyad metal with an atomic weight of 65, combining with numerous acids to form salts. The principal use to which the metal is put is to form the sulphate, from which most of the other salts are made. It is also employed in considerable quantity for the preparation of hydrogen from sulphuric acid: $\text{H}_2\text{SO}_4 + \text{Zn} = \text{ZnSO}_4 + \text{H}_2$. The metal is often contami-

nated with arsenic and in delicate analytical work it is necessary to provide for its freedom from this contamination. The official zinc salts are colorless.

ANALYTICAL REACTIONS.—The alkaline hydroxides give white precipitates in neutral solutions of zinc salts, which precipitates are soluble in excess of the hydroxides.

Soluble carbonates produce white precipitates of zinc carbonate in solutions of zinc salts.

Sulphuretted hydrogen in alkaline solution or ammonium sulphide yields white precipitate of zinc sulphide.

Zinc Sulphate. *Zinci Sulphas.* $\text{ZnSO}_4 \cdot 7\text{H}_2\text{O}$.—*White vitriol*, as this salt is often termed, is the product of the reaction of dilute sulphuric acid upon zinc, as shown above. The solution is evaporated and allowed to crystallize, and is purified by recrystallization. The salt is in colorless prisms or acicular needles, efflorescent in dry air, odorless, nauseous taste, and very soluble in water.

The *impurities* likely to be present in this and the salts made from it are chloride, iron, copper, lead, alkaline earths and alkalies.

Medicinally it is used externally as an astringent antiseptic, and internally as a very prompt emetic, in doses of from 0.50 to 2.00 Gms. In large doses it is poisonous, and care should be exercised not to mistake it for other harmless salts, such as epsom salts, etc.

Zinc Bromide. *Zinci Bromidum.* ZnBr_2 .—Potassium bromide in solution with zinc sulphate produces double decomposition, yielding zinc bromide and potassium sulphate, which both remain in solution, and from which the sulphate is precipitated by the addition of alcohol. The remaining solution of the bromide is evaporated to dryness or the salt

is granulated. Among a number of methods that just described is regarded as the best.

The salt thus obtained is very deliquescent, odorless, of sharp taste, very soluble, and may contain all of the impurities mentioned under sulphate. The medicinal value of the salt is due to the bromine present. It is occasionally prescribed for the relief of insomnia.

Zinc Valerianate. *Zinci Valerianas.* $\text{Zn}(\text{C}_5\text{H}_9\text{O}_2)_2 \cdot 2\text{H}_2\text{O}$.—The zinc sulphate is precipitated by sodium valerianate, and zinc valerianate rises to the surface in the solution, contrary to the usual phenomenon that precipitates deposit: $2\text{Na} \cdot \text{C}_5\text{H}_9\text{O}_2 + \text{ZnSO}_4 = \text{Zn}(\text{C}_5\text{H}_9\text{O}_2)_2 + \text{Na}_2\text{SO}_4$. The salt occurs in white pearly scales, is permanent, slightly soluble in water, and has a faint odor of valerianic acid. It is a form for the administration of valerianic acid and possesses the same medicinal properties, being used as an antispasmodic and nerve tonic in doses of 0.03 to 0.18 Gm.

Precipitated Zinc Carbonate. *Zinci Carbonas Præcipitatus.* $(\text{ZnCO}_3)_2 \cdot 3\text{Zn}(\text{OH})_2$.—Zinc sulphate is precipitated by sodium carbonate and the resulting insoluble carbonate thoroughly washed with hot water: $5\text{ZnSO}_4 + 5\text{Na}_2\text{CO}_3 + 3\text{H}_2\text{O} = (\text{ZnCO}_3)_2 \cdot 3\text{Zn}(\text{OH})_2 + 5\text{Na}_2\text{SO}_4 + 3\text{CO}_2$. The formula will show that the "carbonate," so called, is really a mixture of carbonate and hydroxide. It is an impalpable powder, insoluble except in acids, odorless and tasteless. This carbonate is much finer and purer than the native carbonate, *calamine*, could conveniently be made. Its principal use is for external application, either as powder or in form of ointment.

Zinc Oxide. *Zinci Oxidum.* ZnO .—The carbonate, made as above, when calcined, loses its CO_2 , becoming ZnO —e.g., $\text{ZnCO}_3 + \text{heat} = \text{CO}_2 + \text{ZnO}$. Much of the commercial

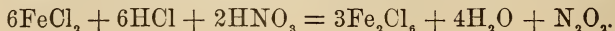
oxide is made from calamine, but for medicinal purposes only the very finest, made from the precipitated carbonate, should be employed. The oxide is a soft, whitish powder, permanent in the air, insoluble and without odor or taste.

Zinc Oxide Ointment.—An ointment made by incorporating 20 per cent of the finest quality of zinc oxide with benzoinated lard.

Zinc Acetate. *Zinci Acetas.* $\text{Zn}(\text{C}_2\text{H}_3\text{O}_2)_2 \cdot 2\text{H}_2\text{O}$.—Easily prepared by adding excess of zinc carbonate to acetic acid, filtering and evaporating the solution or setting it aside to crystallize. $(\text{ZnCO}_3)_3 + 3\text{Zn}(\text{OH})_2 + 10\text{HC}_2\text{H}_3\text{O}_2 = 5\text{Zn}(\text{C}_2\text{H}_3\text{O}_2)_2 + 2\text{CO}_2 + 8\text{H}_2\text{O}$. Zinc acetate is in soft, pearly scales or tablets, efflorescent in dry air; it is soluble in alcohol and water, has a faint odor of acetic acid, and sharp, metallic taste.

Zinc Iodide. *Zinci Iodidum.* ZnI_2 .—Zinc and iodine are allowed to remain in contact in water until solution takes place: $\text{Zn} + \text{I}_2 = \text{ZnI}_2$. The solution is evaporated and the salt obtained in granular form usually. It is a white, granular powder, very deliquescent, without odor. Its use is similar to that of potassium iodide.

Solution of Zinc Chloride. *Liquor Zinci Chloridi.* ZnCl_2 .—Zinc is dissolved in hydrochloric acid; the solution, to which some nitric acid has been added to oxidize any iron present, is then evaporated to dryness and the residue dissolved in water. The second solution is clarified by shaking with zinc carbonate, which precipitates any iron that may be still present, as a carbonate: $\text{Zn} + 2\text{HCl} = \text{ZnCl}_2 + \text{H}_2$. If iron is present it becomes FeCl_2 , thus: $\text{Fe} + 2\text{HCl} = \text{FeCl}_2 + \text{H}_2$. The FeCl_2 , ferrous chloride, is oxidized to the ferric chloride with the nitric acid and more HCl , thus:



The Fe_2Cl_6 when heated probably decomposes into the insoluble ferric oxide and HCl , the latter volatilizing: $\text{Fe}_2\text{Cl}_6 + \text{heat} + 3\text{H}_2\text{O} = \text{Fe}_2\text{O}_3 + 6\text{HCl}$. Any Fe_2Cl_6 remaining unoxidized is removed by the addition of the zinc carbonate, which produces the insoluble ferric hydroxide and carbonate: $\text{Fe}_2\text{Cl}_6 + 3\text{ZnCO}_3 = \text{Fe}_2(\text{CO}_3)_3 + 3\text{ZnCl}_2$. The $\text{Fe}_2(\text{CO}_3)_3$ together with some hydroxide precipitates and the ZnCl_2 remains in solution with the rest of the ZnCl_2 . The $\text{Fe}_2(\text{CO}_3)_3$ usually decomposes with some rapidity into Fe_2O_3 thus: $\text{Fe}_2(\text{CO}_3)_3 = \text{Fe}_2\text{O}_3 + 3\text{CO}_2$. The solution is a clear, colorless liquid, odorless, and is used as a disinfecting and antiseptic solution.

Zinc Chloride. *Zinci Chloridum.* ZnCl_2 .—A very deliquescent salt, rarely used, made by evaporating the solution of zinc chloride. It is used externally as an escharotic, and is often mixed with flour and water to mitigate or regulate its effect.

Zinc Phosphide. *Zinci Phosphidum.* Zn_3P_2 .—Made by conducting vapors of phosphorus in current of hydrogen over heated zinc—a dangerous operation. It occurs in crystalline, brittle fragments, or as grayish powder, insoluble, permanent in air. Zinc phosphide is preferred by many to phosphorus for internal administration. Dose, .003 to .006 Gm.

Aluminum Salts.

Aluminum is one of the most abundant of the metals, and occurs very extensively in rocks and clays. Some compounds are made from cryolite, the fluoride of sodium and aluminum, $6\text{NaF}.\text{Al}_2\text{F}_6$, found abundantly in Greenland. Many of the precious stones are aluminum compounds, the metal itself being of considerable value, owing to the diffi-

culty associated with its separation from its compounds. It is only a little more than $2\frac{1}{2}$ times heavier than water, but yet has all the valuable qualities of iron, with none of its negative properties; it does not tarnish or oxidize in the air, and is only about one-fourth as heavy as iron. Advantage is taken of its lightness in the manufacture of small weights, particularly the grain and smaller divisions of the gramme, and delicate apparatus and articles of ornament. Among the usual methods for isolating it is the reduction of its compounds with metallic sodium or potassium.

It is a trivalent or pseudo-triad (Al_2)^{vi} metal; atomic weight, 27; symbol Al.

ANALYTICAL REACTIONS.—The alkaline hydroxides produce gelatinous white precipitates, $\text{Al}_2(\text{OH})_6$, which are soluble in excess, excepting in ammonia. The alkali carbonates yield a similar precipitate, the CO_2 escaping: $\text{Al}_2(\text{SO}_4)_3 + 3\text{Na}_2\text{CO}_3 + 3\text{H}_2\text{O} = \text{Al}_2(\text{OH})_6 + 3\text{Na}_2\text{SO}_4 + 3\text{CO}_2$. Ammonium sulphide also precipitates the hydroxides with the evolution of hydrogen sulphide: $\text{Al}_2(\text{SO}_4)_3 + 3(\text{NH}_4)_2\text{S} + 6\text{H}_2\text{O} = \text{Al}_2(\text{OH})_6 + 3(\text{NH}_4)_2\text{SO}_4 + 3\text{H}_2\text{S}$.

Alum. *Alumen*. $\text{K}_2\text{Al}_2(\text{SO}_4)_4 \cdot 24\text{H}_2\text{O}$.—This salt is the type of all the alums, explained in a previous chapter. It is made chiefly from the aluminum silicate in alum-clay by appropriate treatment with sulphuric acid, the addition of potassium sulphate, and allowing the two sulphates to crystallize together with 24 molecules of water. If in place of potassium sulphate the ammonium sulphate is employed, *ammonium alum*, $(\text{NH}_4)_2\text{Al}_2(\text{SO}_4)_4 \cdot 24\text{H}_2\text{O}$, is formed.

Alum occurs in large octahedral crystals or as white powder, is odorless, has sweetish, astringent taste, soluble in water, and effloresces slightly in dry air. Iron, zinc or lead

may be present as impurities in any of the aluminum compounds, and may be looked for by the usual tests.

Alum is one of the best styptics and astringents we have.

Dried Alum. *Alumen Exsiccatum*. $K_2Al_2(SO_4)_4$.—The crystallized alum when heated loses its water of crystallization and becomes the official *Alumen Exsiccatum*. Exsiccated means dried. Alum loses about 45 per cent of its weight in the process of drying, and dried alum is more powerful than the hydrous alum in consequence.

The loss in weight which a crystallized salt sustains upon drying is easily found from the molecular weights before and after drying, f.i. alum before drying is $K_2Al_2(SO_4)_4 \cdot 24H_2O$, its molecular weight is the sum of the weights of the atoms:

$$\begin{aligned} 2Al &= 2 \times 27 = 54 \\ 2K &= 2 \times 39 = 78 \\ 4(SO_4)_4 &= 4 \times (S = 32 + 4O = 64) = 96 = 384 \\ 24H_2O &= 24 \times (H_2 = 2 + O = 16) = 432 \end{aligned}$$

$$\text{Molecular weight of alum} = 948$$

Upon drying the $24H_2O$ are expelled; hence there is a loss of 432; then $948 - 432 = 516$, the molecular weight of $K_2Al_2(SO_4)_4$.

To express the loss in percentage it is only necessary to use the rule of three: $948 : 432 :: 100 : x = 45.57$ per cent loss.

100 grains of alum would therefore become dried or *burnt* alum when the weight would have become reduced to 55 grains.

Aluminum Hydrate. *Alumini Hydras*. $Al_2(OH)_6$.—*Hydrox-*

ide would be a better term than hydrate to apply to this drying or absorbing and antacid powder. It is made by precipitating sodium carbonate by alum and thoroughly washing the precipitate with hot water and drying. (See reaction above.) The light white powder thus obtained is permanent in air, odorless and insoluble. It is used but little.

Aluminum Sulphate. *Alumini Sulphas.* $\text{Al}_2(\text{SO}_4)_3 \cdot 18\text{H}_2\text{O}$.—The freshly prepared hydroxide is dissolved in dilute sulphuric acid and the solution allowed to crystallize or granulate. It is antiseptic, and seldom or rarely used internally.

Cerium.

THERE is only one cerium salt official, the *oxalate*, and that is made by a very circumstantial process, necessitated by the association of cerium with the two rare metals lanthanum and didymium. The cerium compounds when sufficiently purified are precipitated by oxalic acid.

Cerium Oxalate. *Cerii Oxalas.* $\text{Ce}_2(\text{C}_2\text{O}_4)_3 \cdot 9\text{H}_2\text{O}$.—Is a whitish powder, slightly granular, permanent in the air, odorless, tasteless, insoluble, soluble in hydrochloric acid. Sodium hypochlorite gives a red precipitate in the latter solution, soluble in warm HCl, evolving chlorine. Oxalate of cerium is regarded as almost a specific for checking nausea. Dose, 0.06 to 0.30 Gm.

Lead Salts.

OF the several combinations in which lead is found in nature, the sulphide, or *galena*, is the one usually chosen as the source of the metal and its salts. Galena occurs very

abundantly in this country and is usually associated with silver, though often found in the free state. The metal is extracted from the ore by roasting first in contact with and in the latter part of the process without, access of air, or by melting with charcoal.

Lead is a bluish-white, soft, malleable metal, with a perceptible taste, and an odor when rubbed, little acted upon by air, but soluble to a slight extent in very pure water, becoming a hydroxide. Water thus impregnated is very poisonous, and increases in its toxic qualities if the hydroxide becomes changed, through the presence of carbon dioxide, into the more soluble bicarbonate. A knowledge of this fact should prompt the utmost carefulness in the use of lead pipe or leaden vessels of any kind for purposes where presence of lead might entail disastrous results. There are numerous cases on record where poisoning was traced to lead in drinking-water or in mineral-waters drawn through leaden pipes. Magnesium sulphate or sodium, or highly diluted sulphuric acid, are chemical antidotes. Lead is a dyad; atomic weight, 206.4; sp. gr., 11.45; symbol, Pb.

ANALYTICAL REACTIONS. — 1. Sulphuretted hydrogen produces a black precipitate of the sulphide, PbS . Very minute quantities may be detected by this test.

2. Any soluble iodide gives a characteristic yellow precipitate of lead iodide, PbI_2 .

3. Soluble sulphates precipitate the insoluble lead sulphate, PbSO_4 .

4. Soluble carbonates precipitate the insoluble lead carbonate, PbCO_3 .

Lead Oxide. *Plumbi Oxidum*. PbO .—Litharge is one of the five oxides of lead, and is made by roasting the lead ores

in a reverberatory furnace. It is often obtained as a by-product in the extraction of silver from argentiferous galena. Litharge is a heavy yellowish or reddish powder, insoluble, colorless, tasteless, remaining unchanged in the air. It may contain as impurities traces of alkalies and alkaline earths, zinc and carbonate. From it most of the other salts of lead are prepared.

Lead Acetate. *Plumbi Acetas.* $\text{Pb}(\text{C}_2\text{H}_3\text{O}_2)_2 \cdot 3\text{H}_2\text{O}$.—This is the most important of the lead salts used in pharmacy; it is easily made by heating together a mixture of lead oxide and acetic acid: $\text{PbO} + 2\text{HC}_2\text{H}_3\text{O}_2 = \text{Pb}(\text{C}_2\text{H}_3\text{O}_2)_2 + 3\text{H}_2\text{O}$. If the acetic acid is not pure the salt usually has a smoky odor which does not disappear rapidly upon recrystallization. The commercial salt should not be used in pharmacy, but only that made from the purest acetic acid. *Sugar of lead*, as the salt is most frequently called, occurs in colorless, shining, prismatic crystals or scales, is soluble in alcohol and water, has faintly acetous odor and astringent taste, and absorbs carbonic acid from the air. The latter is the cause of the turbidity noticed in solutions of the salt. This turbidity may be removed by the addition of a little acetic acid. Acetate of lead is astringent, and is used externally in solution as a sedative wash, often combined with laudanum.

Solution of Lead Subacetate. *Liquor Plumbi Subacetatis.*—This aqueous liquid, containing about 25 per cent of lead subacetate, approximately $\text{Pb}_2\text{O}(\text{C}_2\text{H}_3\text{O}_2)_2$, is obtained by boiling together lead oxide and lead acetate in water previously boiled to expel all its dissolved air. In the operation air should be excluded as much as possible to prevent the formation of carbonate. The salt in solution is not a definite compound, but consists of several oxyacetates in

variable proportions. The reaction between the oxide and acetate is sometimes expressed in this wise: $3\text{Pb}(\text{C}_2\text{H}_3\text{O}_2)_2 + 3\text{PbO} = \text{Pb}_3\text{O}(\text{C}_2\text{H}_3\text{O}_2)_4 \cdot \text{Pb}_3\text{O}_2(\text{C}_2\text{H}_3\text{O}_2)_2$.

GOULARD'S EXTRACT is a term often applied to the solution. It is a clear, colorless liquid of alkaline reaction, usually depositing a sediment after contact with air, which consists of carbonate. The solution is never used internally; but externally it is preferred to a solution of the acetate.

Diluted Solution of Lead Subacetate. *Liquor Plumbi Subacetatis Dilutus.*—LEAD-WATER.—A solution of 3 parts Goulard's extract in 97 of boiled water which has been cooled. The formation of opalescence in the solution is due to the carbonic acid gas in the water. It may be dispelled with a few drops of acetic acid. *Lead-water* has at times been mistaken for *lime-water*, and serious results have attended the mistakes. It has been suggested that the opalescence in lead water be allowed to remain, as a mark of distinction from lime water, which should always be dispensed as a clear and transparent solution. Lead-water is used externally as a cooling and soothing application.

Liniment of Lead Subacetate.—This sedative and cooling application to burns, which is a mixture of 40 parts of solution subacetate lead and 60 parts cottonseed oil, has been dismissed from the 1890 U. S. P.

Cerate of Lead Subacetate. *Ceratum Plumbi Subacetatis.*—GOULARD'S CERATE.—Twenty parts of solution subacetate lead are incorporated with 80 parts of camphor cerate; the resulting preparation has the qualities of the solution in ointment form. The addition of a few drops of acetic acid prevents the cerate from becoming yellow.

Lead-Plaster. *Emplastrum Plumbi.*—*Diachylon* or *lead-plas-*

ter is a chemical compound, an oleate of lead, prepared by boiling together lead oxide and olive oil. Olive oil contains, among other fatty bodies, the oleate of propenyl, $C_3H_5(C_{18}H_{33}O_2)_3$, which reacts upon the PbO thus: $2C_3H_5(C_{18}H_{33}O_2)_3 + 3PbO + 3H_2O = 3Pb(C_{18}H_{33}O_2)_2$ (oleate of lead or lead-plaster) + $2C_3H_5(OH)_3$ (glycerin). The lead-plaster is a by-product in the manufacture of glycerin. Lead-plaster is chiefly valuable as a base for many of the medicinal plasters.

DIACHYLON OINTMENT.—Lead-plaster reduced to the consistency of an ointment with olive oil constitutes the diachylon ointment. It is perfumed with oil of lavender.

Lead Nitrate. *Plumbi Nitras.* $Pb(NO_3)_2$.—Lead nitrate is readily obtained by dissolving lead oxide in nitric acid by means of heat, filtering, evaporating and crystallizing: $PbO + 2HNO_3 = Pb(NO_3)_2 + H_2O$. It is in form of transparent or whitish, nearly opaque crystals, permanent in air, soluble in water, odorless, astringent taste. It is used mainly in making the lead iodide, though at times it is used externally.

Lead Iodide. *Plumbi Iodidum.* PbI_2 .—A solution of lead nitrate is precipitated by a solution of a soluble iodide, potassium iodide usually, and the precipitate washed and dried. $Pb(NO_3)_2 + 2KI = PbI_2 + 2KNO_3$. Lead iodide is a heavy, bright-yellowish powder, unchanged in air, odorless, tasteless, and practically insoluble in alcohol and water, but soluble in solution of ammonium chloride and in solution of the acetates of the alkalies.

The iodide is sometimes given internally in doses of 0.03 to 0.10 Gm., but it is chiefly used in form of ointment.

Ointment of Lead Iodide. *Unguentum Plumbi Iodidi.*—Ten

parts of lead iodide are incorporated with 90 parts of benzoinated lard to obtain the ointment of lead iodide.

Lead Carbonate. *Plumbi Carbonas.* $(\text{PbCO}_3)_2\text{Pb}(\text{OH})_2$.—**WHITE-LEAD.**—The carbonate is made on a very extensive scale for use as a pigment by exposing metallic lead to the action of the air, acetic acid vapors and decaying matter which yields carbonic acid gas. It may be made by precipitating a solution of lead nitrate with sodium carbonate, washing and drying the precipitate: $\text{Pb}(\text{NO}_3)_2 + \text{Na}_2\text{CO}_3 = \text{PbCO}_3 + 2\text{NaNO}_3$. It is not a true carbonate, but contains, as its formula indicates, a certain quantity of hydroxide. Its chief use is in the arts as a base for paints. In pharmacy it is used externally in form of ointment, and very often injudiciously in cosmetics and hair-restorers.

Ointment of Lead Carbonate consists of 10 parts of lead carbonate and 90 of benzoinated lard thoroughly incorporated; it is used externally to reduce inflammation.

Copper Salts.

Copper is found in a native state, but it also occurs as oxide, sulphide, arsenate, carbonate, sulphate, etc. It is of a red color; specific gravity, 8.95 to 9, and is more valuable in the arts and manufactures than in pharmacy. It is a dyad; atomic weight, 63.18; and it forms two official salts, the sulphate and acetate; symbol, Cu.

ANALYTICAL REACTIONS.—1. H_2S produces a black precipitate of the sulphide, CuS .

2. Red cupric ferrocyanide is precipitated from copper solutions by ferrocyanide of potassium.

3. Ammonia-water gives deep-blue color with dilute solutions of copper salts and a blue precipitate in strong solutions.

4. Metallic zinc or iron is coated with copper in solutions of copper, an equivalent quantity of the zinc or iron going into solution.

Copper Sulphate. *Cupri Sulphas.* $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$.—BLUE VITRIOL OR BLUE-STONE.—This salt is made on a large scale by dissolving waste copper in warm dilute sulphuric acid, evaporating, crystallizing and purifying by recrystallization. It is in form of large, translucent, deep-blue crystals, efflorescent, odorless, nauseous taste, soluble in water, insoluble in alcohol. Internally in doses of 0.30 Gm. it is emetic, in 0.03 Gm. doses tonic; externally its solution is a good stimulating wash. It is largely used as a disinfectant together with sulphate of iron.

The impurities sometimes present are iron, alkalies and alkaline earths, lead, zinc and arsenic.

Copper Acetate, $\text{Cu}(\text{C}_2\text{H}_3\text{O}_2)_2 \cdot \text{H}_2\text{O}$.—Of several methods of preparing this salt, that in which lead acetate is employed in solution with copper sulphate is the most convenient one: $\text{CuSO}_4 + \text{Pb}(\text{C}_2\text{H}_3\text{O}_2)_2 = \text{PbSO}_4 + \text{Cu}(\text{C}_2\text{H}_3\text{O}_2)_2$. The lead solution should be concentrated. The solution of the acetate is filtered and allowed to crystallize, when it yields deep-green prismatic crystals which are odorless, efflorescent, soluble in water, and only slightly soluble in alcohol. The salt is sometimes, though erroneously, called *verdigris*. It is used in much the same manner as the sulphate is, but is not any longer official.

Mercury and Its Salts.

Mercury, or quicksilver, is one of the only two liquid elements, bromine being the other, known to chemists. The sulphide, or *cinnabar*, HgS , is its most abundant natural compound, and from it usually by heating it in fur-

naces with access of air, the metal itself is obtained. The metal is volatile, and, becoming disengaged from the sulphur, which burns and escapes as SO_2 in the process, its vapors are conducted into suitable chambers and there condensed : $\text{HgS} + \text{O}_2 = \text{Hg} + \text{SO}_2$. There are other processes of extraction, but they are more complicated and little used. The extensive mines of California yield at present more than enough to supply the markets of this country, and considerable quantities are exported. There are very rich mines in Spain also which have been worked from pre-Christian times, and are still productive.

The mercury salts are a very important class of compounds, pharmaceutically and chemically, but the chief and most extended use of the metal is found in the arts—in the manufacture of thermometers, barometers, mirrors, in mining gold and silver by amalgamation, and in making the pigment *vermilion*, etc. For pharmaceutical purposes the crude mercury is purified, preferably by distillation, or by squeezing through chamois leather. The latter method removes only suspended matter or mechanical impurities, as traces of oxide or water. Mercury is a silver-white, bright, liquid metal, odorless, tasteless, insoluble except in some acids; it is slightly volatile at ordinary temperatures; its volatility increases with the temperature, being greatest at its boiling-point, 662°F ., and least at its freezing-point, -40°F . Mercury sometimes contains tin or other metals as impurities, and when it does, globules dropped upon white paper do not retain their globular form nor roll about freely, and frequently leave traces or streaks on the paper. The metal should have a bright surface and should be dry.

Mercury has an atomic weight of 200 (exactly, 199.80) and forms two series of compounds, the *mercurous* and the *mercuric*. It is a dyad, and in the mercuric salts a single atom has bivalent value, while in the mercurous salts two atoms together have a combining value of two. The two chlorides will illustrate this by their graphic formulæ:

Mercuric Chloride, HgCl_2 .



Mercurous Chloride, Hg_2Cl_2 .



The base in the mercuric salts is always a *single* atom of mercury, Hg , while the base in the mercurous salts always consists of *two* atoms— Hg_2 . In the mercurous compounds one bond from each of two atoms of mercury neutralizes or combines with each other, leaving the one other bond from each atom to unite with a radical equivalent to two bonds.

Important Reactions of Compounds of Mercury.

	With Mercuric Salts in Solution.	With Mercurous Salts.
Hydrochloric acid gives :	No precipitate, HgCl_2 soluble (corrosive sublimate).	White precipitate, in- soluble, Hg_2Cl_2 (cal- omel).
Ammonia - water gives :	White precipitate (call- ed "white precipi- tate") NH_2HgCl .	Black precipitate, $\text{NH}_2\text{Hg}_2\text{Cl}$.
Hydrogen sulphide gives :	Black precipitate, HgS (sulphide).	Black precipitate, Hg_2S (sulphide).
Potassium iodide gives :	Red precipitate, HgI_2 (red mercuric iodide).	Green or yellow pre- cipitate, Hg_2I_2 .
Potassium hydrox- ide gives :	Yellow precipitate, HgO (yellow oxide).	Nearly black precip- itate, Hg_2O .

All the salts of mercury are poisonous, but the mercuric are far more so than the mercurous. There is no specific antidote for salts of mercury, but white of egg and other albuminous substances have proved of some value by coagulating the soluble salts of mercury when administered in time. Free use should be made of emetics or the stomach-pump, and eggs and milk given freely.

A piece of bright copper or a bright penny (made so by dipping into nitric acid) becomes coated with mercury in solutions of mercury salts, an equivalent quantity of copper going into solution.

Mercury dissolved in hot and concentrated nitric acid yields the mercuric nitrate, but if the nitric acid is diluted and cold, the oxidation is incomplete and the mercurous nitrate is formed.

Preparations Containing Metallic Mercury.

No other element is used so largely in the metallic state in medicine and pharmacy as mercury, and the United States Pharmacopœia recognizes five such preparations.

Mass of Mercury. *Massa Hydrargyri*.—BLUE PILL or BLUE MASS.—Mercury is rubbed or triturated with honey of rose and glycerin until globules are no longer visible under a magnifying glass of ten-diameter power—that is, until the metal is thoroughly subdivided and incorporated with the excipients. This work is laborious and tedious when executed with pestle and mortar, and resort is profitably had to machinery, especially when manufacturing it on the large scale, as is done at present. When thorough subdivision has been accomplished, powdered licorice and marsh-mallow are added to make a mass of a pilular consistence.

This preparation contains 33 per cent of metallic mer-

cury, and is usually given in doses up to 0.66 to 1.00 gm. as a purgative in bilious affections, its administration to be followed by a mild saline cathartic.

Mercurial Ointment. *Unguentum Hydrargyri*.—**BLUE OINTMENT.**—Mercury is *extinguished* or *killed*, as the process of subdividing mercury by trituration is often called, with previously prepared mercurial ointment, or, as the 1890 U. S. P. directs, with oleate of mercury, and incorporating with it melted lard and suet. This ointment is a convenient form for the external administration of mercury, and is largely used. It contains 50 per cent of mercury.

Mercurial Plaster. *Emplastrum Hydrargyri*.—This plaster is employed in medicine when it is not convenient to use the ointment. It contains 30 per cent of mercury extinguished with oleate of mercury and incorporated with lead-plaster.

Ammoniac Plaster with Mercury. *Emplastrum Ammoniac cum Hydrargyro*.—This plaster is somewhat milder than the above, containing only 18 per cent of mercury extinguished with oleate of mercury, besides lead-plaster. The ammoniac is previously emulsified with dilute acetic acid.

Mercury with Chalk. *Hydrargyrum cum Creta*. Is a preparation of metallic mercury in powder form, and serves the same purpose as the mass does. Mercury, clarified honey and chalk are triturated until the mercury is no longer visible under a ten-diameter lens, the trituration being facilitated by the addition of a small quantity of water. Lastly enough chalk is added to make the preparation contain 38 per cent strength of mercury, and the whole thoroughly dried. The preparation should be kept in well-stoppered bottles protected from light to prevent the formation of traces of oxide of mercury.

Salts of Mercury.

The salts of mercury should be kept in dark, amber-colored, well-stoppered bottles, protected from light. When handling them it should be borne in mind that nearly all them are very poisonous.

Corrosive Mercuric Chloride. *Hydrargyri Chloridum Corrosivum.* HgCl_2 .—CORROSIVE SUBLIMATE, CORROSIVE CHLORIDE OF MERCURY.—This, one of the most poisonous and valuable of the salts of mercury, is made from mercuric sulphate by subliming it with common salt—e.g., $\text{HgSO}_4 + 2\text{NaCl} = \text{HgCl}_2 + \text{Na}_2\text{SO}_4$.

The HgSO_4 is obtained from metallic mercury by boiling with strong sulphuric acid: $2\text{H}_2\text{SO}_4 + \text{Hg} = \text{HgSO}_4 + \text{SO}_2 + 2\text{H}_2\text{O}$.

The corrosive chloride occurs in crystalline masses (to distinguish it from calomel, which is in powder form), is odorless, has persistent metallic taste, is permanent in air, soluble in water, alcohol and ether, and volatilizes when heated to a high degree, yielding vapors which are extremely poisonous.

Arsenic is sometimes detected as a contamination. Corrosive sublimate is one of the most powerful and extensively employed antiseptics. It has come into wide use of late in antiseptic surgery, proving the most valuable and powerful antiseptic and disinfectant in use. Its other medicinal applications are numerous. Internally it is very valuable as an alterative; its dose ranges from 0.001 to 0.005 Gm. The serious consequences which may attend carelessness in the handling of this salt should prompt extreme caution in its use.

Mild Mercurous Chloride. *Hydrargyri Chloridum Mite.*—SUB-

CHLORIDE OF MERCURY, MILD CHLORIDE OF MERCURY, CALOMEL.—Mercurous sulphate, Hg_2SO_4 , mixed with common salt and heated, produces a sublimate of the mercurous chloride: $\text{Hg}_2\text{SO}_4 + 2\text{NaCl} = \text{Hg}_2\text{Cl}_2 + \text{Na}_2\text{SO}_4$. The mercuric sulphate when rubbed with another atom of mercury becomes the mercurous sulphate: $\text{HgSO}_4 + \text{Hg} = \text{Hg}_2\text{SO}_4$.

The sublimate of calomel is obtained in powder form by rapid and sudden cooling of the vapors; slower cooling yields the sublimate in mass form. Calomel, the *Hydrargyri Chloridum Mite* of the Pharmacopœia, is in form of a white, heavy, impalpable powder (*Hydrargyri Chloridum Corrosivum* is kept in form of crystalline masses), and is odorless, tasteless, wholly insoluble in ordinary solvents, and quickly blackened by ammonia-water. It is permanent in the air, but volatilizes at high temperatures; it may be reduced to metallic mercury by heating with dried sodium carbonate in a long test-tube.

On account of the similarity in the processes for making both the mild and corrosive chloride, the former may contain traces of the latter, a possibility which must be strenuously guarded against. Mercuric chloride may be detected in calomel by washing the latter with alcohol or ether, which dissolves out any corrosive chloride that might be present, and testing the filtrate for chloride with solution silver nitrate, and for mercury with H_2S . Calomel is totally insoluble in ordinary solvents, and water or alcohol, after having been agitated with it, should remain unaffected by any reagent.

AMMONIATED MERCURY, which resembles calomel in appearance, is soluble in acetic acid, from which solution it may be precipitated as the black sulphide with H_2S gas.

The last three tests should be applied before putting the calomel upon the dispensing-counter.

Owing to its insolubility in the system its action is largely mechanical, and its dose comparatively large, ranging from .030 to 1.00 Gm. In small doses, from .006 to .030 Gm., it is alterative and stimulating to the liver; in large doses purgative. It was at one time believed that sodium chloride converted calomel partly into corrosive sublimate, but extensive investigation and trials have failed to verify this. The pharmacist should nevertheless be vigilant to prevent its being dispensed with salts concerning whose chemical action toward it there may be a doubt. Calomel enters into the Compound Antimony Pills, and into the Compound Cathartic Pills.

Red Mercuric Iodide. *Hydrargyri Iodidum Rubrum.* HgI_2 .—RED MERCURIC IODIDE, BINIODIDE OF MERCURY.—Obtained by adding potassium iodide to mercuric chloride in solution, and thoroughly washing and drying the precipitate at a low temperature. The decomposition is represented by the following equation: $\text{HgCl}_2 + 2\text{KI} = \text{HgI}_2 + 2\text{KCl}$. As the red iodide is soluble in excess of potassium iodide solution, care should be taken to avoid an excess. It is a scarlet-red crystalline powder, odorless, tasteless, almost insoluble in water, slightly soluble in alcohol and permanent in the air. Water shaken with the salt should not be affected by solution of nitrate of silver—test for the absence of soluble iodide or chloride.

Mercuric iodide in medicine is given internally chiefly as an alterative in syphilis; dose, .002 to .015 Gm.

Yellow Mercurous Iodide. *Hydrargyri Iodidum Flavum.* Hg_2I_2 . GREEN MERCUROUS IODIDE, PROTOIODIDE OF MERCURY.—Iodine and mercury are rubbed together in equivalent

quantities in presence of alcohol, and when combination is completed washed with alcohol. The alcohol prevents by its evaporation the elevation of temperature which the chemical union induces; alcohol is employed to wash the finished salt because it dissolves out the traces of mercuric iodide which are usually formed, and which are much more poisonous than the mercurous iodide: $\text{Hg}_2 + \text{I}_2 = \text{Hg}_2\text{I}_2$. The U. S. P. directs the salt to be made by precipitating mercurous nitrate solution by potassium iodide solution: $\text{Hg}_2(\text{NO}_3)_2 + 2\text{KI} = \text{Hg}_2\text{I}_2 + 2\text{KNO}_3$.

The green or yellow iodide is so called because it has a greenish-yellow color; it becomes quite yellow when exposed to air, wherefore it is termed *yellow iodide* in the U. S. P. It is practically insoluble in water, insoluble in alcohol, odorless, tasteless, and becomes darker in color if exposed to light for a long time.

The medical properties of the yellow iodide are the same as those of the red, but because of its milder action it is sometimes preferred to the latter salt. The dose is somewhat larger, being often given in doses of .065 Gm.

Yellow Mercuric Oxide. *Hydrargyri Oxidum Flavum.* HgO .—Made by adding solution of mercuric chloride to solution of soda and setting aside for several hours, when the precipitate formed is thoroughly washed with distilled water and dried between blotting-paper: $\text{HgCl}_2 + 2\text{NaOH} = \text{HgO} + 2\text{NaCl} + \text{H}_2\text{O}$.

Yellow mercuric oxide is a light, orange-yellow powder, becoming darker when exposed to light, insoluble in water, tasteless and odorless. It has the same composition as the red oxide. It is chiefly used for the ointment and for the oleate of mercury.

Ointment of Yellow Mercuric Oxide. *Unguentum Hydrargyri*

Oxidi Flavi.—Contains 10 parts of the yellow oxide thoroughly incorporated with 90 of simple ointment. It and the ointment of red oxide mercury are used medicinally in similar indications.

Oleate of Mercury. *Oleatum Hydrargyri*.—Made by dissolving 20 parts of yellow mercuric oxide in 80 of oleic acid. Chemical combination takes place as follows: $\text{HgO} + 2\text{HC}_{18}\text{H}_{33}\text{O}_2$ (oleic acid) = $\text{Hg}(\text{C}_{18}\text{H}_{33}\text{O}_2)_2$ (oleate mercury) + H_2O . Heat should not be employed. The oleate is employed when it is not convenient to use the ointment.

When kept for some time metallic mercury separates, hence the preparation should never be kept long.

Red Mercuric Oxide. *Hydrargyri Oxidum Rubrum*. HgO .—**MERCURIC OXIDE, RED PRECIPITATE**.—This oxide differs from the yellow probably only in color; chemically it is the same. It may be made by decomposing mercuric nitrate with heat, e.g.: $2\text{Hg}_2(\text{NO}_3)_2 + \text{heat} = 2\text{HgO} + 4\text{NO}_2 + \text{O}_2$.

The salt is usually in heavy, orange-red crystalline powder, sometimes in scales, odorless, tasteless, insoluble in alcohol and water, soluble in nitric or hydrochloric acid. It differs from the yellow oxide in not forming a mercuric oxalate of a white color, as that does, when digested on water bath with solution of oxalic acid.

Ointment of Red Mercuric Oxide. *Unguentum Hydrargyri Oxidi Rubri*.—This ointment is the principal form in which the red oxide is used. It contains 10 per cent red oxide rubbed smooth with castor oil and incorporated with simple ointment; it is much employed by ophthalmologists and dermatologists, and is popularly known as red precipitate ointment. It should not be dispensed except upon a physician's order.

Mercuric Cyanide. *Hydrargyri Cyanidum*. HgCy_2 or $\text{Hg}(\text{CN})_2$.—This salt of mercury is prepared by passing hydrocyanic

acid gas into mercuric oxide: $\text{HgO} + 2\text{HCN} = \text{Hg}(\text{CN})_2 + \text{H}_2\text{O}$. The hydrocyanic acid is made in the process by the action of sulphuric acid upon ferrocyanide of potassium.

The cyanide occurs as colorless or white prismatic crystals darkening on exposure to light; it is odorless, of a bitter taste, soluble in alcohol and water, and exceedingly poisonous. It should be tested for mercuric chloride with iodide of potassium. It is alterative; dose, .003 to .005 Gm.

Ammoniated Mercury. *Hydrargyrum Ammoniatum*. HgNH_2Cl .—WHITE PRECIPITATE, MERCUR-AMMONIUM CHLORIDE.—Obtained by adding solution of mercuric chloride to ammonia-water and washing and drying the precipitate: $2\text{NH}_4\text{HO} + \text{HgCl}_2 = \text{NH}_4\text{Cl} + \text{HgNH}_2\text{Cl} + 2\text{H}_2\text{O}$. There are numerous mercur-ammonium compounds, but the official one probably is as above. White precipitate occurs in form of white powder, odorless, tasteless, insoluble in alcohol or water, and permanent in air. It is very poisonous, and care should be taken not to mistake it for calomel, which resembles it in appearance and physical properties.

Ointment of Ammoniated Mercury. *Unguentum Hydrargyri Ammoniatum*.—Ammoniated mercury is never used internally, and only seldom externally in form of ointment. The ointment contains 10 parts ammoniated mercury and 90 of benzoinated lard, and is employed in various skin-diseases.

Yellow Mercuric Sulphate. *Hydrargyri Subsulphas Flavus*. $\text{Hg}(\text{HgO})_2\text{SO}_4$ or $\text{HgSO}_4 \cdot 2\text{HgO}$.—TURPETH MINERAL, BASIC SULPHATE MERCURY.—Mercuric sulphate prepared as above shown is added to boiling water, when the yellow basic sulphate separates as a precipitate, while some acid mercuric sulphate which is formed remains in solution.

Turpeth mineral is a heavy, lemon-yellow powder, odorless, tasteless, insoluble in water and permanent in air. It

should be free from mercurous salt. If this is absent the salt is soluble in 20 parts of hydrochloric acid. The salt is rarely used—occasionally it is given as an emetic, in doses of .065 to .195 Gm., to children suffering from croup.

Red Mercuric Sulphide. *Hydrargyri Sulphidum Rubrum.* HgS .—**MERCURIC SULPHIDE, CINNABAR, VERMILION.**—Mercury and sulphur are fused and the mass powdered and sublimed. Much of the finest vermilion in commerce comes from China, though some is made in this country. It is a brilliant, dark-red or scarlet powder, odorless, insoluble, and permanent in the air. It is used most extensively in the arts as a pigment; in medicine it is used but little, and is not any longer official.

Solution of Mercuric Nitrate. *Liquor Hydrargyri Nitratis.*—A liquid containing 60 per cent of mercuric nitrate, $\text{Hg}(\text{NO}_3)_2$, together with about 11 per cent of free nitric acid. It is made by dissolving red oxide of mercury in diluted nitric acid. It is a clear, colorless liquid of a specific gravity of 2.10, miscible with alcohol or water. The solution is very corrosive and poisonous, and is employed principally in the treatment of ulcers and malignant growths as a cauterant.

Citrine Ointment of Mercuric Nitrate. *Unguentum Hydrargyri Nitratis.*—**CITRINE OINTMENT.**—Made by thoroughly incorporating solution of mercuric nitrate with a mixture of lard oil and nitric acid. It is employed externally in skin diseases and in some diseases of the eyes.

The Silver Salts.

SILVER occurs chiefly in the metallic state in nature, but its sulphide is also found abundantly either alone or associated with sulphide of lead. The chloride, *horn-silver*, is also found in large enough quantities to make it profitable to extract the silver. Silver is of a brilliant whiteness, very malleable and ductile, and used in pharmacy principally in making the nitrate, from which, in turn, all the other official salts are prepared. It has a specific gravity of 10.5, and is applied to a great variety of purposes in the arts. Coin-silver contains copper to give it greater hardness. Silver is a monad and has an atomic weight of 107.66; symbol, Ag (*argentum*). Metallic silver in form of silver-leaf is used by pharmacists for coating pills.

The silver salts should be kept in well-stoppered bottles protected from light.

ANALYTICAL REACTIONS.—1. Silver is soluble in nitric acid, forming the nitrate, with the evolution of N_2O_2 vapors.

2. Chlorides in solution produce a curdy white precipitate of silver chloride, which is insoluble in nitric acid, but readily soluble in excess of ammonia-water.

3. The alkalis give a brownish precipitate of silver oxide. This precipitate should be handled carefully, as it easily explodes when rubbed in contact with readily oxidizable substances.

4. Metallic silver is blackened by hydrogen or ammonium sulphide, and solutions of silver salts give black precipitates with the same reagents.

5. Solutions of silver salts in contact with the skin turn the latter black, owing to the reduction of silver to the metallic state.

Silver Nitrate. *Argenti Nitras.* AgNO_3 .—This, the most valuable of the silver salts, is made by dissolving pure silver (not coin-silver, which contains copper) in nitric acid, evaporating to dryness, and increasing the heat until the mass melts to expel all nitric vapors. When cooled the mass is dissolved in water and crystals allowed to form in the filtered solution. Any copper which may have been present is thus oxidized and remains on the filter. $3\text{Ag}_2 + 8\text{HNO}_3 = 6\text{AgNO}_3 + \text{N}_2\text{O}_2 + 4\text{H}_2\text{O}$. The nitrate is in form of heavy tubular, transparent, colorless fusible crystals, soluble in water, odorless, of metallic taste, and becoming gray or black upon exposure to light in contact with organic matter. *Metallic impurities* are detected by removing all the silver with hydrochloric acid and evaporating the filtrate to dryness—a fixed residue would indicate the presence of other metals. Sodium nitrate is sometimes thus detected, and if present is usually an adulteration. Copper, if present, gives blue solution with ammonia-water. Silver nitrate is used externally as a caustic and escharotic. Great care should attend its internal administration, as it is very poisonous.

ANTIDOTE.—Common salt produces the insoluble silver chloride: $\text{AgNO}_3 + \text{NaCl} = \text{AgCl} + \text{NaNO}_3$. Emetics should follow.

Fused Silver Nitrate. *Argenti Nitras Fusus.* $\text{AgNO}_3 + \text{AgCl}$. **LUNAR CAUSTIC.**—Made by melting silver nitrate crystals at as low a temperature as possible and adding 4 per cent of hydrochloric acid, heating until the nitric vapors are dispelled, and casting into moulds. The HCl produces AgCl , which, though insoluble, renders the nitrate much less brittle and more tenacious. The fused or moulded silver nitrate is in form of hard, white, solid pencils, having the same proper-

ties that the crystals have, excepting that it is not wholly soluble in water, the chloride, amounting to about 5 per cent., remaining undissolved.

Diluted Silver Nitrate. *Argenti Nitras Dilutus.* $\text{AgNO}_3 + \text{KNO}_3$.—Diluted or *mitigated* silver nitrate is made by melting together equal parts of silver and potassium nitrates in a porcelain crucible and pouring into moulds.

This preparation is used whenever a milder caustic than the pure silver nitrate is desired. Its properties are the same as those of the undiluted nitrate, only less powerful.

Silver Cyanide. *Argenti Cyanidum.* AgCN .—Obtained by adding solution of potassium cyanide to solution of nitrate of silver: $\text{AgNO}_3 + \text{KCN} = \text{AgCN} + \text{KNO}_3$. The potassium nitrate formed is removed by filtration and the washing of the precipitate.

The salt is a white, insoluble powder, permanent in dry air, and becoming brown or black by exposure to light and organic matter. It may also be made by passing hydrocyanic gas into solution of silver nitrate.

Its sole use is in the extemporaneous preparation of hydrocyanic acid, which is made by bringing together prescribed quantities of silver cyanide, hydrochloric acid and water, allowing the precipitated silver chloride to subside, when the solution of the acid is poured off. The latter constitutes the official hydrocyanic acid, which contains 2 per cent. HCN . This acid is classed among the organic acids.

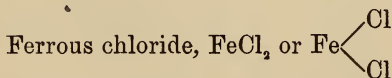
Silver Oxide. *Argenti Oxidum.* Ag_2O .—A solution of silver nitrate is treated with caustic potassa solution, and the precipitate thoroughly washed and dried: $2\text{AgNO}_3 + 2\text{KOH} = \text{Ag}_2\text{O} + 2\text{KNO}_3 + \text{H}_2\text{O}$. Silver oxide is a heavy,

nearly black, and almost insoluble powder. When triturated with organic or easily oxidizable bodies it is liable to cause explosion, owing to the readiness with which it parts with its oxygen. It should not be brought in contact with ammonia. Pills of the oxide frequently explode.

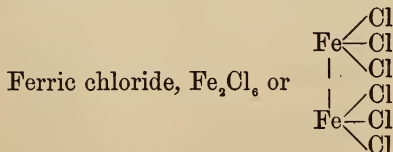
Silver Iodide. *Argenti Iodidum.* AgI .—Potassium iodide and silver nitrate produce silver iodide and potassium nitrate. $\text{AgNO}_3 + \text{KI} = \text{AgI} + \text{KNO}_3$. The iodide is a light, yellow powder, insoluble, odorless and tasteless. If pure it is not changed by light, but it generally turns greenish-yellow in time.

Iron and its Salts.

Iron furnishes a greater number of official salts and preparations than any other metal, and its salts are among the most important in pharmacy and medicine. It is hard, tenacious, ductile, malleable, grayish white, specific gravity 7.76, and oxidizes in moist air. Iron can be made to burn in oxygen. It unites with most of the metals and non-metals, forming two classes of compounds, the *ferrous* and the *ferric*. In the former it has two bonds and in the latter four but only three are capable of uniting with bonds of other elements. For illustration :



The iron has two bonds.



Here one atom of iron has four bonds, but one from each of two atoms combines with or neutralizes each other, so that *two atoms together have six bonds* which will unite with the bonds of other elements. Atomic weight, 56. Symbol, Fe (*ferrum*).

Sources.—Iron occurs very abundantly and is widely distributed in nature. Native iron is only found in meteors or shooting stars. The chief sources of iron are ferric oxide, Fe_2O_3 , and magnetic oxide, Fe_3O_4 (or $\text{FeO} \cdot \text{Fe}_2\text{O}_3$), both of which occur in a variety of forms; the carbonate, FeCO_3 , and the sulphides, of which there are also a great variety, from the FeS to the Fe_7S_8 . The sulphate and phosphate are also found.

Important Analytical Reactions of Iron Salts.

Reagent.	With Ferrous Salts.	With Ferric Salts.
Ferrocyanide potassium, K_4FeCy_6 .	Potassium ferrous-ferrocyanide, $\text{K}_2\text{FeFeCy}_6$. Nearly white, turning blue in air. Everitt's salt.	Ferric ferrocyanide precipitate, $\text{Fe}_4(\text{FeCy}_6)_3$. Deep blue color. Prussian blue.
Ferricyanide potassium, K_3FeCy_6 , or $\text{K}_6\text{Fe}_2\text{Cy}_{12}$.	Ferrous ferricyanide, $\text{Fe}_3(\text{FeCy}_6)_2$. Dark blue. Turnbull's blue. (This is a very valuable test for detecting ferrous in ferric salts.)	No precipitate, but greenish color. $\text{Fe}_2(\text{FeCy}_6)_2$.
Ammonia-water, NH_4OH .	Ferrous hydroxide, $\text{Fe}(\text{OH})_2$. White, turning through green, etc., to brown, becoming ferric.	Ferric hydroxide, $\text{Fe}_2(\text{OH})_6$, brown magma.
Ammonium sulphide, $(\text{NH}_4)_2\text{S}$.	Ferrous sulphide, FeS . Black precipitate.	Ferrous sulphide, FeS , black precipitate. The ferric becomes reduced to ferrous salt.
Sulphocyanide potassium, KCys .	No change.	Intensely red solution of ferric sulphocyanide, $\text{Fe}_2(\text{Cys})_6$.
Tannic acid.	No change at once.	Blue-black precipitate of ferric tannate.

In the study of the long list of iron salts it is again convenient to follow the order of derivation, as given below:

FIRST SERIES.

Iron, metallic.

Iodide, saccharated.

Iodide, syrup.

Lactate, salt.

Chloride, salt.

Chloride, solution.

Chloride, tincture.

Acetate iron and ammonium.

Sulphate, crystals.

SECOND SERIES.

Sulphate, precipitated.

Sulphate, dried.

Pills aloes and iron.

Carbonate, saccharated.

Carbonate, mass.

Carbonate, pills.

Reduced iron.

Iodide pills.

Compound iron mixture.

Hypophosphite, salt.

Subsulphate, solution.

Tersulphate, solution.

THIRD SERIES.

Sulphate iron and ammonium, double salt.

Valerianate, salt.

Hydrated oxide of iron and magnesia.

Hydrated oxide, magma.

FOURTH SERIES.

Hydrated oxide, dried, troches.

Hydrated oxide, dried, plaster.

Acetate, solution.

Nitrate, solution.

Tartrate iron and ammonium, scales.

Tartrate iron and potassium, scales.

Citrate, solution.

Citrate, scales.

Phosphate, soluble, scales.

Phosphate iron, quinine and strychnine,
syrup.

Pyrophosphate, soluble, salts.

Citrate iron and quinine, scales.

Citrate iron and quinine, scales, soluble.

Citrate iron and ammonium, scales.

Citrate, wine.

Citrate iron and quinine, wine.

Citrate iron and strychnine, scales.

By reference to the above arrangement it will be seen that metallic iron furnishes directly or indirectly by derivation all the salts and preparations of iron mentioned in the U. S. P. The salts or preparations, or both, in each series are made directly from the salt or preparation lowest in the column above. Thus, for illustration, the solution of the nitrate is made from the hydrated oxide, which is prepared from the solution of the tersulphate; the latter in turn is derived from the crystals of the sulphate, and the sulphate from metallic iron. Or, in other words, among the series made directly from iron, the sulphate is the starting-point for a second series, of which the

solution of the tersulphate yields a third series, among which is the hydrated oxide, from which the fourth series of salts and preparations, including all the scale salts, is made.

Iron.—The United States Pharmacopœia directs that metallic iron should be in the form of fine, bright and non-elastic wire, that form being usually very pure.

Saccharated Ferrous Iodide. *Ferri Iodidum Saccharatum.* FeI_2 .—Iron and iodine are made to react upon each other in presence of water: $\text{Fe}_2 + 2\text{I}_2 = 2\text{FeI}_2$. The green solution of the ferrous iodide is filtered into sugar of milk, and the mixture evaporated to dryness, triturated with more sugar of milk and a little reduced iron, and afterward preserved in small well-stoppered bottles in a dark and cool place.

The object of adding the sugar of milk is to preserve the iodide, which is ordinarily very unstable, decomposing easily and becoming colored by the free iodine. The reduced iron aids in preventing the separation of iodine. Sugar is employed for its preservative properties in a number of other iron preparations, as will presently be seen. The saccharated iodide of iron is a yellowish-white, sometimes grayish-white powder, dissolving to a clear solution in water. A brown color would indicate the presence of free iodine and preclude its employment for dispensing purposes. Free iodine when present may easily be detected by adding a few drops of gelatinized starch to the aqueous solution—a blue color would indicate its presence.

Medicinally the saccharated iodide is used as an alterative and tonic. Dose, 0.30 gm. It contains 20 per cent FeI_2 .

Syrup of Ferrous Iodide. *Syrupus Ferri Iodidi.* FeI_2 .—The syrup is made as the above, excepting that instead of adding to the solution of ferrous iodide sugar of milk and

evaporating to dryness, simple syrup is added, yielding a preparation having all the properties of the dry saccharated iodide in a liquid state. The syrup should be kept exposed to daylight to aid the sugar in preventing the separation of iodine. If iodine does separate, it combines with the water in presence of daylight to form the colorless hydriodic acid, thereby avoiding the formation of a brown color. Syrup of ferrous iodide has a greenish color, and should not be dispensed if of a brownish color. Its dose is 10 to 15 drops, largely diluted with water, to prevent its action upon the enamel of the teeth, which it invariably destroys when in contact for even a short time. The syrup contains 10 per cent. ferrous iodide; specific gravity about 1.353.

Ferrous Lactate. *Ferri Lactas.* $\text{Fe}(\text{C}_3\text{H}_5\text{O}_3)_2 \cdot 3\text{H}_2\text{O}$.—This ferrous salt results when iron is dissolved in lactic acid: $\text{Fe}_2 + 4\text{HC}_3\text{H}_5\text{O}_3 = 2\text{Fe}(\text{C}_3\text{H}_5\text{O}_3)_2 + 2\text{H}_2$. Whenever metallic iron is dissolved in an acid hydrogen is evolved. The solution of the lactate is evaporated and allowed to crystallize, when the salt is obtained as a pale greenish-white crystalline mass, or, if granulated, in grains. It is soluble in about 40 parts of water, is odorless and has a mild ferruginous taste. Medicinally it is used by some in preference to the other forms of iron, as it is believed to be more readily assimilated. Dose, 0.06 to 0.180 Gm.

Ferric Chloride. *Ferri Chloridum.* $\text{Fe}_2\text{Cl}_6 \cdot 12\text{H}_2\text{O}$.—Obtained by dissolving iron in hydrochloric acid and oxidizing the ferrous chloride so produced to the ferric state with nitric acid, evaporating to dryness, expelling all acid vapors and crystallizing.

When iron is acted upon by HCl , the ferrous chloride is produced: $\text{Fe}_2 + 4\text{HCl} = \text{FeCl}_2 + 2\text{H}_2$. The latter is oxi-

dized to ferric chloride with nitric acid in the presence of more hydrochloric acid. The oxidation may be thus expressed: $6\text{FeCl}_2 + 2\text{HNO}_3 + 6\text{HCl} = 3\text{Fe}_2\text{Cl}_6 + \text{N}_2\text{O}_2 + 4\text{H}_2\text{O}$. Here again it will be well to remember that when nitric acid acts as an oxidizing agent, *two* molecules of it split up into H_2O , N_2O_2 , and 3O , and that the oxygen removes in this case the hydrogen from hydrochloric acid, leaving the chlorine free to unite with the ferrous chloride. The three atoms of oxygen require six of hydrogen to form water, hence the necessity of taking *six* molecules of HCl ; the six free hydrogen atoms require six molecules of FeCl_2 to convert them into Fe_2Cl_6 , because one molecule of Fe_2Cl_6 requires two more atoms of Cl than two molecules of FeCl_2 or Fe_2Cl_4 . If two atoms of chlorine oxidize two molecules of FeCl_2 into Fe_2Cl_6 , the six free hydrogen atoms will require three times two FeCl_2 , or *six* FeCl_2 . The N_2O_2 escapes in form of reddish fumes, and the one molecule of water from the 2HNO_3 added to the three obtained from the six hydrogen atoms from six HCl and the three oxygen atoms from two HNO_3 give the four molecules of water, balancing the equation.

Ferric chloride is in yellow crystalline pieces, very deliquescent and hence very soluble in water, also in alcohol and ether. It has an acid reaction, owing to presence of hydrochloric and sometimes nitric acid. Zinc and copper and the fixed alkalies are sometimes present as impurities, together with ferrous salt and oxychloride. Test solution of ferricyanide of potassium does not give a blue color in absence of ferrous chloride. This salt is rarely used internally, the tincture being preferred. Externally it is used as a styptic.

Solution of Ferric Chloride. *Liquor Ferri Chloridi.* Fe_2Cl_6 .—

If in making the preceding the solution of Fe_2Cl_6 is not evaporated, but so adjusted by the addition of water that it contain 37.8 per cent of the anhydrous ferric chloride, the solution of ferric chloride will be the result. It is a reddish-brown liquid, of strongly acid reaction, styptic taste, specific gravity 1.387, miscible with water, alcohol and ether in all proportions. It is used principally to prepare the extensively used tincture chloride iron.

Tincture of Ferric Chloride. *Tinctura Ferri Chloridi.* Fe_2Cl_6 .—This preparation is made by diluting solution of ferric chloride with alcohol and allowing to stand for at least three months before using. It is made so long before using to permit the formation by the action of the free acid present upon the alcohol of certain compound ethers, among which the ethyl chloride predominates. These ethers are believed to have diuretic properties and a peculiar influence upon the urinary passages. Sometimes a brownish precipitate of ferric oxychloride forms upon standing or upon dilution. This is due to not employing enough hydrochloric acid in the preparation of the solution. The tincture is a clear, brownish liquid of an ethereal odor, styptic taste and acid reaction, and contains about 13.6 per cent of the anhydrous salt. It is the most important liquid preparation of iron. Dose, from 5 to 30 drops, largely diluted with water.

Solution of Iron and Ammonium Acetate. *Liquor Ferri et Ammonii Acetatis.* $\text{Fe}_2(\text{C}_2\text{H}_3\text{O}_2)_6 \cdot \text{NH}_4\text{Cl}$.—BASHAM'S MIXTURE.—Tincture of chloride of iron is added to solution of ammonium acetate containing excess of acetic acid; mutual decomposition takes place according to the reaction, $\text{Fe}_2\text{Cl}_6 + 6\text{NH}_4\text{C}_2\text{H}_3\text{O}_2 = \text{Fe}_2(\text{C}_2\text{H}_3\text{O}_2)_6 + 6\text{NH}_4\text{Cl}$. The solution is flavored with aromatic elixir and glycerin

and water added. This solution should be freshly prepared when needed.

The not unpleasant taste and its agreeable appearance, as well as its acceptability to the stomach, recommend it to a wider employment. Its dose is from 15 to 30 cc.

Ferrous Sulphate. *Ferri Sulphas.* $\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$.—GREEN VITRIOL.—This is the last in the series of salts made directly from iron; it is obtained on the large scale by dissolving old scrap-iron in waste sulphuric acid, and purifying as far as possible by repeated crystallization. This form is usually employed for disinfecting and for other purposes which are not pharmaceutical. For pharmaceutical purposes it should be made from the purest acid and iron wire. The following equation expresses the reaction: $2\text{Fe}_2 + 2\text{H}_2\text{SO}_4 = 2\text{FeSO}_4 + 2\text{H}_2$. Ferrous sulphate remains in solution, from which it is obtained by evaporation and crystallization.

The salt occurs as large pale-green monoclinic prisms, efflorescing and becoming oxidized on exposure to air. It is odorless, has acid reaction, and is soluble in water but insoluble in alcohol. The crystals when heated melt in their water of crystallization. The impure salt, known in commerce as *copperas*, though it contains no copper, is used for disinfecting purposes.

Granulated Ferrous Sulphate. *Ferri Sulphas Granulatus.* $\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$.—The crystals of sulphate of iron are dissolved in water containing a little sulphuric acid and the solution poured into alcohol. The sulphate is insoluble in even dilute alcohol and precipitates in form of a granular powder; some of the impurities present in the crystals are soluble in alcohol, hence this method of precipitation yields conveniently not only a granular salt, but also a

very pure one. It is a pale-green powder less liable to oxidize in the air than the crystals and better adapted for dispensing purposes than the latter. Otherwise it is identical with the crystals.

Dried Ferrous Sulphate. *Ferri Sulphas Exsiccatus.* $\text{FeSO}_4 \cdot \text{H}_2\text{O}$.—A convenient quantity of the powdered crystals of sulphate of iron is exposed to heat until it ceases to lose weight, ascertained by weighing at intervals of 30 to 40 minutes. This process yields a form of sulphate of iron containing only 1 molecule of water of crystallization, 6 being dispelled by the heat. It is a grayish-white powder and is used for making pills, or whenever the crystals are unfit for use on account of either their bulk or water of crystallization. The dried sulphate represents nearly twice as much anhydrous salts as the crystals do.

Pills Aloes and Iron. *Pilulæ Aloes et Ferri.*—These pills contain each 1 grain of dried sulphate of iron and 1 grain of purified aloes, besides the excipients, aromatic powder and confection rose.

Saccharated Ferrous Carbonate. *Ferri Carbonas Saccharatus.* FeCO_3 .—Sulphate of iron crystals are made to react upon sodium bicarbonate and the precipitate of ferrous carbonate thoroughly washed. The drained precipitate is mixed with finely powdered sugar and the whole dried on a water-bath: $\text{FeSO}_4 + 2\text{NaHCO}_3 = \text{FeCO}_3 + \text{Na}_2\text{SO}_4 + \text{CO}_2 + \text{H}_2\text{O}$. Here again the sugar acts as a preserving agent, preventing the decomposition and oxidation of the carbonate, which ordinarily is not stable. It is a grayish powder, odorless, partially soluble in water, and wholly soluble in hydrochloric acid with copious evolution of CO_2 gas. If the precipitate is not thoroughly washed, traces of sulphate will remain. These can be detected by test solu-

tion of chloride of barium, which gives white precipitate in presence of sulphates.

The carbonate is a ferruginous tonic; dose, 0.180 to 1.33 Gm.

Mass of Ferrous Carbonate. *Massa Ferri Carbonatis.* FeCO_3 .—This preparation was formerly known as *Pilula Ferri Carbonatis* because it is usually employed in making pills, its consistence fitting it especially for that purpose. Sulphate of iron and sodium carbonate react upon each other, thus: $\text{FeSO}_4 + \text{Na}_2\text{CO}_3 = \text{Na}_2\text{SO}_4 + \text{FeCO}_3$. The precipitate is thoroughly washed and incorporated with sugar and honey to form a mass of a pilular consistence. Mass of carbonate of iron is popularly known as *Vallet's Mass.* Dose, 0.33 Gm.

Reduced Iron. *Ferrum Reductum.* Fe_2 .—Subcarbonate or hydroxide of iron is heated or calcined until it is reduced to the oxide Fe_2O_3 and this is then reduced to the metallic condition by passing hydrogen gas over it. The reactions $\text{Fe}_2(\text{OH})_6 + \text{heat} = \text{Fe}_2\text{O}_3 + 3\text{H}_2\text{O}$, $\text{Fe}_2\text{O}_3 + 3\text{H}_2 = \text{Fe}_2 + 3\text{H}_2\text{O}$ explain the process.

It is a very fine black powder, wholly soluble in sulphuric acid, and should contain not less than 80 per cent of metallic iron. It will burn in air, becoming the ferric oxide: $2\text{Fe}_2 + 3\text{O}_2 = 2\text{Fe}_2\text{O}_3$, the chemical process being the same as *rusting*, only much more rapid and illustrative of the chemical affinity between the iron and the oxygen of the air. The finely divided iron exhibits a greater surface to the action of the oxygen, and the chemical action is consequently so intense as to produce heat and flame.

Reduced iron is one of the most valued of the chalybeate tonics; it is given in doses of 0.120 to 0.330 Gm. in pill form.

Pills of Ferrous Iodide. *Pilulæ Ferri Iodidi.* FeI_2 .—These pills are made by acting upon reduced iron with iodine, and adding to the filtered solution of the resulting iodide of iron, powdered sugar, licorice root, licorice extract and acacia, and evaporating the mixture until it attains a pilular consistence, when it is divided into pills. To prevent the iodine from becoming liberated and the iron from oxidizing, the Pharmacopœia directs that the pills should be coated with tolu balsam, which is dissolved for the purpose in stronger ether. The pills are rolled in this solution and allowed to dry. The ether evaporates very quickly, leaving a thin, impervious and air-tight coating of balsam upon the pills.

Compound Iron Mixture. *Mistura Ferri Composita.* $\text{FeCO}_3 + \text{K}_2\text{SO}_4$.—Potassium carbonate is mixed with powdered myrrh, sugar, spirit of lavender to flavor, and rose-water, to which mixture, contained in a bottle in which the preparation is intended to be kept, coarsely powdered sulphate of iron is added, and the whole well shaken: $\text{FeSO}_4 + \text{K}_2\text{CO}_3 = \text{FeCO}_3 + \text{K}_2\text{SO}_4$. The mixture should be freshly prepared when wanted. It contains the sulphate of potassium, which is formed, in addition to the carbonate, wherein it differs from the mass of carbonate of iron and the pills of iodide of iron, in each of which the formed sulphate is washed out and does not remain in the preparation. The mixture is similar in this respect to the compound pills of iron. It is also known as Griffith's Mixture.

Pills of Carbonate of Iron. *Pilulæ Ferri Carbonatis.* $\text{FeCO}_3 + \text{K}_2\text{SO}_4$.—Potassium carbonate is rubbed with powdered myrrh and then with sulphate of iron, and when the reaction has ceased the mass is made into pills with

tragacanth, althæa, glycerin and water: $\text{FeSO}_4 + \text{Na}_2\text{SO}_4 = \text{FeCO}_3 + \text{Na}_2\text{SO}_4$. The reaction sometimes does not set in until the pills become softened in the stomach.

Ferric Hypophosphite. *Ferri Hypophosphis.* $\text{Fe}_2(\text{H}_2\text{PO}_2)_6$.—Made by double decomposition between calcium hypophosphite and ferrous sulphate: $\text{FeSO}_4 + \text{Ca}(\text{H}_2\text{PO}_2)_2 = \text{Fe}(\text{H}_2\text{PO}_2)_2$ (ferrous hypophosphite) + CaSO_4 . The calcium sulphate precipitates, while the ferrous hypophosphite is held in solution. The application of heat during the evaporating of the solution changes the ferrous salt to the ferric condition. The ferric salt is a white or grayish powder, permanent in the air, odorless, nearly tasteless, and only slightly soluble in water. Ferric phosphate should be absent; if it is not, the salt is not completely soluble in acetic acid.

Hypophosphite of iron was at one time highly recommended in the treatment of consumption, but its efficiency was probably overestimated. It is now given as a blood and nerve tonic in pill or syrup form. Dose up to 0.66 Gm.

Solution of Ferric Subsulphate. *Liquor Ferri Subsulphatis.* $\text{Fe}_4\text{O}(\text{SO}_4)_6$.—SOLUTION OF BASIC FERRIC SULPHATE, MONSEL'S SOLUTION.—Sulphate of iron is added to a boiling mixture of sulphuric and nitric acids until effervescence ceases. The following equation approximately expresses the reaction: $12\text{FeSO}_4 + 3\text{H}_2\text{SO}_4 + 4\text{HNO}_3 = 3\text{Fe}_4\text{O}(\text{SO}_4)_6 + 5\text{H}_2\text{O} + 2\text{N}_2\text{O}_2$. The effervescence is due to the disengagement of the N_2O_2 vapors, which, as soon as they are generated, unite with two atoms of oxygen, $\text{N}_2\text{O}_2 + \text{O}_2 = \text{N}_2\text{O}_4$, becoming the tetroxide of nitrogen, whose fumes, and not those of N_2O_2 , have a reddish color. Enough distilled water is added to adjust the percentage of metallic iron to 13.6.

Monssel's solution is a dark, reddish-brown, heavy liquid of a specific gravity of 1.55. It is nearly odorless, has a very astringent taste, and is miscible in all proportions with water and alcohol. Ferrous salt, which is sometimes present, is detected with test-solution of ferricyanide of potassium, with which it gives a deep blue color. This solution when diluted is the most valuable preparation for stopping bleeding or external hemorrhages. It is employed internally in doses of from 1 to 5 minims, well diluted. In some medical works, and among some physicians, the solution of *subsulphate* of iron is known as solution of *persulphate* of iron. Whenever the latter is prescribed the former should be dispensed, according to the U. S. Pharmacopœia.

Solution of Ferric Sulphate. *Liquor Ferri Tersulphatis.* $\text{Fe}_2(\text{SO}_4)_3$.—SOLUTION OF NORMAL FERRIC SULPHATE.—The method for making this solution is identical with that for the preparation of the preceding, except that the quantity of H_2SO_4 is increased sufficiently to convert *all* of the ferrous sulphate employed into the ferric condition: $12\text{FeSO}_4 + 6\text{H}_2\text{SO}_4 + 4\text{HNO}_3 = 6\text{Fe}_2(\text{SO}_4)_3 + 8\text{H}_2\text{O} + 2\text{N}_2\text{O}_2$.

In comparing the equation expressing the reaction in the preparation of the solution of subsulphate of iron with that for the solution of tersulphate of iron, it will be seen that in the former 3 molecules of H_2SO_4 are employed, and in the latter 6, or just twice as many. This excess of H_2SO_4 in the latter over the former is just sufficient to convert the basic ferric sulphate into the normal ferric sulphate: $3\text{Fe}_4\text{O}(\text{SO}_4)_5 + 3\text{H}_2\text{SO}_4 = 6\text{Fe}_2(\text{SO}_4)_3 + 3\text{H}_2\text{O}$. The strength of the solution in ferric sulphate is adjusted with distilled water to 28.7 per cent. Solution of tersulphate of iron has nearly the same properties as the solution of the subsulphate has, differing only in containing

more sulphuric acid and having a lower specific gravity—1.32. The solution of the subsulphate, when two volumes of it are slowly mixed with one volume of H_2SO_4 , deposits a white solid mass on standing, while the tersulphate solution does not. This solution is used principally in the preparation of the third series of the iron preparations.

Ferric Ammonium Sulphate. *Ferri et Ammonii Sulphas.* $\text{Fe}_2(\text{NH}_4)_2(\text{SO}_4)_4 \cdot 24\text{H}_2\text{O}$. — AMMONIO FERRIC SULPHATE, AMMONIA FERRIC ALUM, IRON ALUM.—Made simply by dissolving crystals of ammonium sulphate in solution tersulphate of iron and allowing the double salt to crystallize. The salt is in form of pale-violet, eight-sided crystals, efflorescent on exposure to air, sometimes melting in warm weather in its water of crystallization. It is odorless, styptic and soluble in water, insoluble in alcohol. In form of saturated or strong solution it is employed as a styptic. It is one of the alums. (See Alums.)

Ferric Valerianate. *Ferri Valerianas.* $\text{Fe}_2(\text{C}_5\text{H}_9\text{O}_2)_6$.—Solution of tersulphate of iron is precipitated with a solution of sodium valerianate, the precipitated valerianate of iron collected, thoroughly washed and dried.

It is a dark, tile-red, amorphous powder, having a faint odor of valerianic acid; insoluble in water, soluble in alcohol, and decomposed by boiling water, the valerianic acid being set free and ferric hydroxide remaining. The salt is very little used in pharmacy.

Ferric Hydrate with Magnesia. *Ferri Oxidum Hydratum cum Magnesia.* $\text{Fe}_2(\text{OH})_6 + \text{MgSO}_4$. — Calcined magnesia is rubbed to a smooth and thin mixture with water and added to tersulphate of iron solution and the whole well shaken together. The preparation is official because it is antidotal to arsenous acid. It should be freshly prepared

when wanted; the magnesia should be kept ready suspended in water, next to a bottle containing the prescribed amount of the diluted iron solution, so that the preparation can be made at a moment's notice. When the two are mixed, the magnesia being in excess, ferric hydroxide and magnesium sulphate are produced: $\text{Fe}_2(\text{SO}_4)_3 + 3\text{MgO} + 3\text{H}_2\text{O} = \text{Fe}_2(\text{OH})_6 + \text{MgSO}_4$. The ferric hydroxide becomes in a short time converted into the oxyhydrate, $\text{Fe}_2\text{O}_2(\text{OH})_2$, by losing water, thus: $\text{Fe}_2(\text{OH})_6 - 2\text{H}_2\text{O} = \text{Fe}_2\text{O}_2(\text{OH})_2$; a long exposure converts it finally into the ferric oxide, Fe_2O_3 , by losing another molecule of water: $\text{Fe}_2\text{O}_2(\text{OH})_2 - \text{H}_2\text{O} = \text{Fe}_2\text{O}_3$. In the proportion in which the ferric hydroxide loses water, it becomes worthless as an antidote to arsenic, hence the great importance of making the antidote when needed. The ferric hydroxide becomes deoxidized by the arsenic and becomes the insoluble ferrous arsenate and ferrous hydroxide, according to this equation: $2\text{Fe}_2(\text{OH})_6 + \text{As}_2\text{O}_3$ (arsenous acid) = $\text{Fe}_3(\text{As}_2\text{O}_4)_2$ (ferrous arsenate) + $\text{Fe}(\text{OH})_2$ (ferrous hydroxide) + $5\text{H}_2\text{O}$. The sulphate of magnesium formed aids, through its cathartic action, to remove the insoluble arsenate of iron, but it is better to employ prompt emetics shortly after the administration of the antidote.

Ferric Hydrate. *Ferri Oxidum Hydratum.* $\text{Fe}_2(\text{OH})_6$.—Instead of using, as in the preceding, the oxide of magnesium, the hydroxide of ammonium is used as the precipitant in this preparation. It is, however, better to pour the solution of tersulphate of iron into the ammonia-water, which has been previously diluted: $\text{Fe}_2(\text{SO}_4)_3 + 6\text{NH}_4\text{OH} = \text{Fe}_2(\text{OH})_6 + 3(\text{NH}_4)_2\text{SO}_4$. The whole is poured upon a wet muslin strainer and thoroughly washed with water. The product is the same as in the preceding, but without the presence

of any other salt, since the ammonium sulphate is washed out. The ferric hydroxide thus prepared is as efficient an antidote to arsenic as the above, the latter having the advantage, though, of not being over-acid nor over-alkaline, the magnesia being very slightly alkaline and always in excess. The U. S. P. directs that the ingredients for making ferric hydroxide should always be kept on hand, ready for immediate use. The hydroxide freshly prepared is a brownish-red magma gradually turning into the oxide Fe_2O_3 upon exposure or upon drying. It is not used medicinally otherwise than stated; pharmaceutically it is of importance, inasmuch as from it the fourth series of iron preparations including the important scale salts are made.

Troches of Iron. *Trochisci Ferri.* Fe_2O_3 .—These troches are made by massing dried hydroxide of iron with sugar, vanilla and mucilage of tragacanth, and cutting into convenient forms, so that each troche contains 0.33 Gm. of the dried hydroxide of iron.

Iron-Plaster. *Emplastrum Ferri.*—The ordinary strengthening plaster contains dried hydroxide of iron incorporated with olive oil, Burgundy pitch and lead-plaster by means of gentle heat.

Solution of Ferric Acetate. *Liquor Ferri Acetatis.* $\text{Fe}_2(\text{C}_2\text{H}_3\text{O}_2)_6$.—Ferric hydroxide, freshly prepared from solution tersulphate of iron and ammonia-water and thoroughly washed, is dissolved in glacial acetic acid and enough water added to the solution to make it contain 31 per cent. of an hydroferric acetate: $\text{Fe}_2(\text{OH})_6 + 6\text{HC}_2\text{H}_3\text{O}_2 = \text{Fe}_2(\text{C}_2\text{H}_3\text{O}_2)_6 + 6\text{H}_2\text{O}$. The solution is a dark reddish-brown clear liquid with a pleasant acetous odor and sweetish taste; specific gravity, 1.16. Ferrous salt may be present and may be found as usual with test-solution ferricyanide of potas-

sium. This preparation is used principally in the preparation of the official tincture.

Tincture Acetate of Iron. *Tinctura Ferri Acetatis.* $\text{Fe}_2(\text{C}_2\text{H}_3\text{O}_2)_6 + \text{C}_2\text{H}_5\text{C}_2\text{H}_3\text{O}_2$.—Made from the solution of acetate of iron by diluting with alcohol and adding acetic ether, $\text{C}_2\text{H}_5\text{C}_2\text{H}_3\text{O}_2$. This tincture is preferred by some physicians, especially by the German, to the tincture of chloride of iron. It decomposes in time, light and heat causing it to deposit a brown precipitate of probably an oxide. It should therefore be kept in a cool and dark place.

Solution of Ferric Nitrate. *Liquor Ferri Nitratis.* $\text{Fe}_2(\text{NO}_3)_6$.—Solution of ferric nitrate is obtained by dissolving freshly-prepared ferric hydroxide in nitric acid and adding enough water to make the solution contain six per cent of normal ferric nitrate. Thus prepared, solution of ferric nitrate is a transparent, reddish liquid having a strongly styptic taste and acid reaction; specific gravity, 1.05. Medicinally the preparation is an astringent tonic, and is given in doses up to 10 minims, 0.66 cc.

Scale Salts.

The class of compounds known as the Scale Salts of Iron is a series of iron preparations made by dissolving freshly precipitated ferric hydroxide in tartaric or citric acid or in their acid salts, and evaporating the solutions at a *low heat* until of a syrupy consistency, then spreading on plates of glass or porcelain, and allowing to dry thereon until the compounds separate in scales. The secondary products remain and form part of the compounds, which become, according to the acid or acid salt used, potassio-citrate of iron or potassio-tartrate of iron, or similar ammonium or sodium compounds of iron. They are of varying chemical

constitution, and the formulas and equations following are probably only approximate, yet they aid to illustrate the composition of the compounds. The separate processes of the official scale salts are described below.

Iron and Potassium Tartrate. *Ferri et Potassii Tartras.*
 —POTASSIO-FERRIC TARTRATE.—Probably $\text{Fe}_2\text{K}_6(\text{C}_4\text{H}_4\text{O}_6)_6$. To some fresh ferric hydroxide a quantity of acid potassium tartrate is added, and after solution has been effected by means of a water bath, in which the temperature does not rise above 140°F. (or 60°C.), the whole is filtered and evaporated to a syrupy consistency, and poured on glass plates to dry and scale. The addition of a little ammonia during evaporation prevents the formation of a precipitate after solution. The reaction may be according to the equation $\text{Fe}_2(\text{OH})_6 + 6\text{KHC}_4\text{H}_4\text{O}_6 = \text{Fe}_2\text{K}_6(\text{C}_4\text{H}_4\text{O}_6)_6 + 6\text{H}_2\text{O}$. This and the succeeding compound have a chemical similarity to Rochelle salt, the double tartrate of sodium and potassium— $\text{KNaC}_4\text{H}_4\text{O}_6$. In the former the hydrogen of the acid tartrate of potassium has been replaced by ferric iron, Fe_2 , and in the latter by sodium. The normal tartrate of iron is only very slightly soluble.

This compound is in transparent garnet-red scales, odorless, sweetish, soluble in water but not in alcohol, and is used as a mild iron tonic in doses up to 1.00 Gm.

Iron and Ammonium Tartrate. *Ferri et Ammonii Tartras.*—AMMONIO-FERRIC TARTRATE.—Probably $\text{Fe}_2(\text{NH}_4)_6(\text{C}_4\text{H}_4\text{O}_6)_6$. Prepared as the above, substituting acid tartrate of ammonium for the potassium salt. It possesses the same properties as the preceding salt does, only being sometimes of a yellowish-brown color. The acid ammonium salt is made as part of the process by adding carbonate of ammo-

nium to tartaric acid in molecular proportions: $(\text{NH}_4)_2\text{CO}_3 + 2\text{H}_2\text{C}_4\text{H}_4\text{O}_6 = 2\text{NH}_4\text{HC}_4\text{H}_4\text{O}_6 + \text{CO}_2 + \text{H}_2\text{O}$.

Solution of Ferric Citrate. *Liquor Ferri Citratis.* $\text{Fe}_2(\text{C}_6\text{H}_5\text{O}_7)_2$.—Ferric hydroxide is dissolved in citric acid and the solution adjusted to a strength of 35.5 per cent with water. $\text{Fe}_2(\text{OH})_6 + 2\text{H}_3\text{C}_6\text{H}_5\text{O}_7 = \text{Fe}_2(\text{C}_6\text{H}_5\text{O}_7)_2 + 6\text{H}_2\text{O}$. It is a dark-brown liquid, odorless, and has a ferruginous taste; specific gravity, 1.25. Owing to the fact that it is just half the strength of the scales, the pharmacist will find it very convenient to use when the citrate is prescribed in aqueous solution.

Soluble Ferric Phosphate. *Ferri Phosphas Solubilis.*—Phosphate of sodium is dissolved in a solution of ferric citrate and the solution scaled. The scales are a sodio-ferric-citrophosphate. This, like the other scale salts, is not a definite compound, and its formula can only be given approximately. This scale compound contains Na_2HPO_4 and $\text{Fe}_2(\text{C}_6\text{H}_5\text{O}_7)_2$ in a state of combination. It is in bright-green scales, odorless, very soluble in water and insoluble in alcohol. Its advantage over the normal phosphate is in its solubility. Dose up to 0.66 Gm.

Syrup of the Phosphates of Iron, Quinine and Strychnine. *Syrupus Ferri, Quininæ et Strychninæ Phosphatum.*—Quinine sulphate and strychnine are dissolved in solution of phosphate of iron containing an excess of phosphoric acid, and sugar and distilled water added to make a syrup. The excess of phosphoric acid dissolves the alkaloids, converting them into phosphates. The syrup contains .02 per cent. of strychnine alkaloid and 3.00 per cent of quinine sulphate.

Soluble Ferric Pyrophosphate. *Ferri Pyrophosphas Solubilis.*—**FERRIC PYROPHOSPHATE.** Made as the preceding, employ-

ing the pyrophosphate in place of the phosphate of sodium. It occurs in thin green, transparent scales, and is very freely soluble in water, on account of which it is frequently employed in solution. Dose up to 5 grains, 0.33 gm.

Ferric Citrate. *Ferri Citras.* $\text{Fe}_2(\text{C}_6\text{H}_5\text{O}_7)_2$.—The scales of citrate of iron are obtained by simply evaporating the solution of citrate of iron. They are of a garnet-red color, are odorless, and slowly but wholly soluble in water and insoluble in alcohol. Because the salt is very slowly soluble it is not so frequently used as the succeeding one. Dose up to 1.33 Gm.

Iron and Ammonium Citrate. *Ferri et Ammonii Citras.*—AMMONIO-FERRIC CITRATE. Made, as the above, by the addition of a little water of ammonia to the solution of citrate of iron before evaporating. The addition of ammonia furnishes a salt which is more soluble than the above but otherwise identical with it.

Wine of Ferric Citrate. *Vinum Ferri Citratis.*—An agreeable chalybeate tonic prepared by dissolving citrate of iron and ammonium in white wine and adding syrup and tincture sweet orange-peel to flavor. It contains 4 per cent of citrate of iron and ammonium. Dose, up to 4 cc. Sometimes called *sweet wine of iron*.

Bitter Wine of Iron. *Vinum Ferri Amarum.*—Made by dissolving soluble iron and quinine citrate in white wine and adding tincture of sweet orange-peel and syrup. It contains 5 per cent. of the scale salt.

Iron and Strychnine Citrate. *Ferri et Strychninæ Citras.*—Strychnine is dissolved in a solution of citrate of iron and ammonium containing citric acid, and the solution evaporated and scaled. Thus obtained the compound is in form of transparent garnet-red scales, deliquescent on exposure

to air, odorless, bitter taste, wholly soluble in water and only slightly soluble in alcohol. It contains one per cent of strychnine. Dose up to 0.20 Gm.

Soluble Iron and Quinine Citrate. *Ferri et Quininae Citras Solubilis.*—Quinine and citric acid are dissolved in solution of citrate of iron, and the whole evaporated and scaled. It forms thin, transparent scales of a greenish, golden-yellow color. Slowly deliquescent upon exposure to the air, odorless, bitter taste, very soluble in water and but slightly soluble in alcohol. It should contain 12 per cent. of quinine, but very often it is found in the market with a deficiency of quinine.

The Manganese Salts.

The salts of manganese are not of much importance in pharmacy and medicine; there are only two official, the dioxide and the sulphate. Manganese is found in nature mainly as the dioxide, MnO_2 , called pyrolusite or braunstein; the Mn_2O_3 and Mn_3O_4 and the carbonate also occur in nature.

Manganese is present in the radical of potassium permanganate.

Manganese Dioxide. *Mangani Dioxidum.* MnO_2 .—Black Oxide of Manganese.—The U. S. Pharmacopœia recognizes the native crude product, which contains at least 66 per cent of the pure dioxide. It is a heavy, grayish-black gritty powder, odorless, tasteless, and permanent in the air, insoluble. Its chief use is in the preparation of oxygen; when heated it gives off oxygen, becoming the Mn_3O_4 — $3\text{MnO}_2 + \text{heat} = \text{Mn}_3\text{O}_4 + \text{O}_2$.

Manganese Sulphate. *Mangani Sulphas.* $\text{MnSO}_4 + 4\text{H}_2\text{O}$.—When the dioxide is digested with concentrated sulphuric

acid the sulphate is formed with the liberation of oxygen :
 $2\text{MnO}_2 + 2\text{H}_2\text{SO}_4 = 2\text{MnSO}_4 + 2\text{H}_2\text{O} + \text{O}_2$. This salt occurs in colorless or pale rose-colored transparent prisms which are somewhat efflorescent in dry air, very soluble and of a bitter and astringent taste.

The aqueous solution of the salt yields with ammonium-sulphide test-solution a flesh-colored precipitate, soluble in dilute acids.

If a little of the dry salt is mixed with a little sodium hydroxide and the mixture fused, it will yield a dark green mass which dissolves in water with a green color.

The Arsenic Salts.

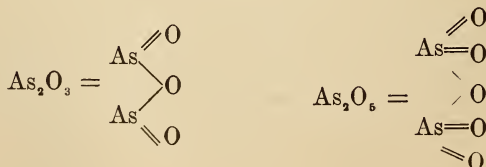
Occurrence.—The element arsenic is sometimes found in a native state, but the largest proportion is derived from the sulphides, *realgar*, As_2S_2 , and *orpiment*, As_2S_3 , and from the arsenical iron ores, the latter consisting chiefly of FeAs_2 , NiAs_2 and CoAs_2 . Mispickel, $\text{Fe}_3\text{S}_2\text{As}$, is also a source, and arsenous acid, As_2O_3 , is found in nature as the mineral *arsenolite*. There are many other forms of occurrence, but they are not utilized much as sources of the element.

Preparation.—The ores, preferably the arsenical iron ores, are roasted in a reverberatory furnace and the volatile product, crude arsenous oxide, As_2O_3 , condensed in long and nearly horizontal chimneys, or in a series of chambers made of brick. The arsenic volatilizes and unites with oxygen, becoming the arsenous oxide, $\text{As}_4 + 3\text{O}_2 = 2\text{As}_2\text{O}_3$, which condenses and forms the crude oxide which is purified by resublimation. The purified As_2O_3 is then heated in retorts with charcoal, which unites with the oxygen,

leaving the arsenic in the elementary state as a sublimate:
 $2\text{As}_2\text{O}_3 + 3\text{C}_2 = \text{As}_4 + 6\text{CO}$ (or 3CO_2).

Properties, Physical and Chemical.—Arsenic is a steel-gray, crystalline solid, having a lustre, and a specific gravity ranging from 5.70 to 5.95. It is brittle, tarnishes in air but not in pure water, for which reason it is often kept in distilled water. When heated it volatilizes, but heated in air it becomes oxidized and sublimes as As_2O_3 . It is both triad and pentad, more often the former; its atomic weight is 75 (exactly 74.9) and its vapor density (compared to hydrogen) 149.8, which is twice its atomic weight, so that its molecule in the gaseous state occupies only half as much space as the molecule of hydrogen does. Arsenic is therefore, like phosphorus, an exception to the rule that the atomic weights of elements under similar conditions give equal volumes of vapor, or, in other words, that the vapor density of an element is half its molecular weight, a molecule being supposed to consist of two atoms.

It appears from this that the molecule of arsenic consists of four atoms, As_4 . Arsenic also, like iron, mercury and some of the other metals, forms two series of compounds the *arsenous* and the *arsenic*, exemplified in the oxides As_2O_3 , arsenous, and As_2O_5 , arsenic. In the former it has 3 bonds, in the latter 5:



Arsenic is considered by some chemists as being a non-

metal, but the majority class it as a *metalloid* (*oid* means like) because its chemical behavior is similar to that of the metals in many respects. It behaves in the dual capacity, chemically, of a non-metal and of a metal. In the former it resembles phosphorus, uniting, e.g., with oxygen to form oxides, which oxides with water produce the corresponding acids, from which corresponding salts may be prepared. The following will illustrate:

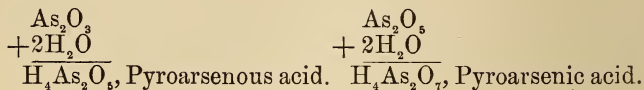
Oxides	As_2O_3	Arsenous Oxide.	As_2O_5	Arsenic Oxide.
	$+ 3\text{H}_2\text{O}$	Water.	$+ 3\text{H}_2\text{O}$	
Acids	$2\text{H}_3\text{AsO}_3$	Arsenous Acid.	$2\text{H}_3\text{AsO}_4$	Arsenic Acid.
-	$3\text{H}_2 + 3\text{K}_2$	Replacing H by K.	$-3\text{H}_2 + 3\text{K}_2$	Replacing H by K.
Salts	$2\text{K}_3\text{AsO}_3$	Arsenite of Potassium.	$2\text{K}_3\text{AsO}_4$	Arsenate of Potassium.

Oxides	P_2O_3	Phosphorous Oxide.	P_2O_5	Phosphoric Oxide.
	$+ 3\text{H}_2\text{O}$		$+ 3\text{H}_2\text{O}$	
Acids	$2\text{H}_3\text{PO}_3$	Phosphorous Acid.	$2\text{H}_3\text{PO}_4$	Phosphoric Acid.
-	$3\text{H}_2 + 3\text{K}_2$	Replacing H by K.	$-3\text{H}_2 + 3\text{K}_2$	Replacing H by K.
Salts	$2\text{K}_3\text{PO}_3$	Phosphite of Potassium.	$2\text{K}_3\text{PO}_4$	Phosphate of Potassium.

The above are all normal or ortho-compounds. A further similarity between phosphorus and arsenic will be seen in the following comparison of their meta- and pyro-compounds:

Oxides	As_2O_3		As_2O_5	
	$+ \text{H}_2\text{O}$		$+ \text{H}_2\text{O}$	
Acids	2HAsO_2	Metarsenous Acid.	2HAsO_3	Metaarsenic Acid.
-	$\text{H}_2 + \text{K}_2$		$-\text{H}_2 + \text{K}_2$	
Salts	2KAsO_2	Metarsenite of Potassium. (Fowler's Solution.)	KAsO_3	Metarsenate of Potassium.
Oxides	P_2O_3		P_2O_5	
	$+ \text{H}_2\text{O}$		$+ \text{H}_2\text{O}$	
Acids	2HPO_2	Metaphosphorous Acid.	2HPO_3	Metaphosphoric Acid.
-	$\text{H}_2 + \text{K}_2$		$-\text{H}_2 + \text{K}_2$	
Salts	2KPO_2	Metaphosphite of Potassium.	KPO_3	Metaphosphate of Potassium.

In a like manner the pyroacids and pyrosalts may be obtained by adding two molecules of water to each of the oxides of arsenic and phosphorus. Any other metal may be employed to replace the potassium, if in equivalent proportions.



The salts ortho-, pyro-, and meta-arsenites and arsenates are therefore obtained, in theory, by replacing the hydrogen in the acids as above shown—the arsenic in them behaves as a non-metal, resembling phosphorus. In the other capacity arsenic acts as a metal; that is, it occupies the base in a number of salts, in such as AsCl_3 , AsBr_3 , As_2S_3 , etc. (See Antimony.)

Important Analytical Reactions.—1. HYDROGEN SULPHIDE produces in acid solutions precipitate of the *yellow* sulphide, $\text{As}_2\text{O}_3 + 3\text{H}_2\text{O} + 3\text{H}_2\text{S} = \text{As}_2\text{S}_3 + 6\text{H}_2\text{O}$.

2. ARSENIC SALTS are soluble in ammonium sulphide: $\text{As}_2\text{S}_3 + 3(\text{NH}_4)_2\text{S} = 2(\text{NH}_4)_3\text{As}_2\text{S}_3$.

3. MARSH'S TEST consists in reducing arsenic compounds to metallic arsenic by means of nascent hydrogen. With the latter arsenic enters into combination to form *arsenous hydride* or *arsenuretted hydrogen*, which is identified by numerous reactions, but usually by its behavior when passed into silver nitrate solution, or when ignited and its flame directed upon cold porcelain. The hydrogen is usually generated from H_2SO_4 by means of zinc: $2\text{H}_2\text{SO}_4 + \text{Zn} = 2\text{ZnSO}_4 + 2\text{H}_2$. The H in contact with arsenic becomes AsH_3 thus: $\text{As}_2\text{O}_3 + 6\text{H}_2 = 2\text{H}_3\text{As} + 3\text{H}_2\text{O}$. The H_3As is a very *poisonous gas*, and care should be had not to inhale

it. It burns with a yellow flame, producing As_2O_3 thus: $2\text{H}_3\text{As} + 3\text{O}_2 = 3\text{H}_2\text{O} + \text{As}_2\text{O}_3$. If the flame be directed against *cold* porcelain or glass, the reduction of temperature prevents the oxidation of the arsenic, and it deposits in black or steel-gray spots: $4\text{H}_3\text{As} + 3\text{O}_2 = 6\text{H}_2\text{O} + \text{As}_4$. Antimonuretted hydrogen produces similar spots under similar conditions. (See Antimony).

The arsenic spots are dissolved by a solution of hypochlorite of sodium: $\text{As}_4 + 10\text{NaClO} + 6\text{H}_2\text{O} = 4\text{H}_3\text{AsO}_4 + 10\text{NaCl}$. (Difference from antimony spots, which remain unaffected.)

If the H_3As is passed into solution of silver nitrate, the silver is reduced to metallic condition and precipitates: $\text{H}_3\text{As} + 6\text{AgNO}_3 + 3\text{H}_2\text{O} = 3\text{Ag}_2 + \text{H}_3\text{AsO}_3 + 6\text{HNO}_3$. This test is usually employed to distinguish arsenic from antimony—the antimony precipitates, the arsenic does not.

4. FLEITMAN'S TEST consists in generating hydrogen from metallic zinc or aluminum and strong solution of soda or potassa by heating. The addition of the solution of an arsenic salt will produce arsenuretted hydrogen, which may be tested as above: $2\text{Al}_2 + 4\text{KOH} + \text{As}_2\text{O}_3 + \text{H}_2\text{O} = 2\text{H}_3\text{As} + 2\text{K}_2\text{Al}_2\text{O}_4$. The value of this method is that it does not produce antimonuretted hydrogen, affording a ready means of distinguishing between arsenic and antimony salts.

5. REINSCH'S TEST is applied by adding a piece of bright copper to an (HCl) acidulated solution of an arsenic salt and boiling. If arsenic is present, the copper will become coated with a deposition of metallic arsenic. It is essential that the copper be chemically pure.

6. BERZELIUS' TEST. When the arsenic oxides are heated in a long test-tube with charcoal, the arsenic becomes re-

duced to the metallic condition and deposits on the inside of the tube as an iron-gray mirror. Arsenic in state of vapor has a *garlicky* odor, which may be noticed in the carrying out of this test. The vapor is poisonous. $4\text{As}_2\text{O}_3 + 3\text{C}_2 = 2\text{As}_4 + 6\text{CO}_2$.

7. AMMONIACAL SOLUTION OF COPPER SULPHATE gives a grass-green precipitate in solutions of arsenic. The precipitate dissolves in NH_4OH . It consists of $\text{Cu}_3\text{As}_2\text{O}_6 \cdot 2\text{H}_2\text{O}$ (or $3\text{CuO} \cdot \text{As}_2\text{O}_3 \cdot 2\text{H}_2\text{O}$) and is known as Scheele's green.

8. AMMONIO-NITRATE OF SILVER SOLUTION produces a lemon-yellow precipitate of arsenite of silver in solutions of arsenic, which dissolves in an excess of ammonia. (Ammonio-nitrate of silver solution is prepared by adding NH_4OH to AgNO_3 until the precipitate at first produced is nearly all dissolved.) $6\text{NH}_4\text{OH} + 6\text{AgNO}_3 + 2\text{As}_2\text{O}_3 = 2\text{Ag}_3\text{AsO}_3 + 6\text{NH}_4\text{NO}_3 + 3\text{H}_2\text{O}$.

Arsenous Oxide. *Acidum Arsenosum.* As_2O_3 .—WHITE ARSENIC.—This compound is popularly known as *arsenic* and as *arsenous acid*. Both terms are incorrect, but they have been so thoroughly established by usage that the Pharmacopœia has not discarded their use.

Arsenous acid consists, as shown above, of H_3AsO_3 ; all acids contain hydrogen. The As_2O_3 is obtained, as above shown, in the preparation of metallic arsenic. From it all other pharmacopœial compounds are made. It is a heavy, white, opaque powder or semi-transparent mass, permanent in the air, volatile when heated, odorless, tasteless, slowly soluble in water, freely in HCl and in the alkaline hydroxides and carbonates. Its strength of purity should be at least 98.80 per cent.

Arsenous acid is very poisonous and is often bought with suicidal intention; the careful pharmacist will only dispense

it upon physicians' orders. Its antidote, as for all arsenic compounds, is freshly precipitated ferric hydroxide (see Ferric Hydroxide), which forms an insoluble, and hence harmless, salt with it: $2\text{Fe}_2(\text{OH})_6 + \text{As}_2\text{O}_3 = \text{Fe}_3(\text{AsO}_4)_2$ (ferrous arsenate) + $\text{Fe}(\text{OH})_2$ (ferrous hydroxide) + $5\text{H}_2\text{O}$.

The oxide is given in doses up to 0.003 Gm. Medicinally it is an alterative.

Solution of Arsenous Acid. *Liquor Acidi Arsenosi.* H_3AsO_3 .—This preparation is a solution of the oxide As_2O_3 in water, the solution being aided by heat and the addition of a small quantity of hydrochloric acid. The latter does not enter into combination with the oxide; it merely facilitates solution.

The solution is a convenient means for the administration of arsenic in liquid form, but it is less popular with physicians than the following. Dose is up to 5 minims, 0.33 cc. It contains 1 per cent of As_2O_3 .

Solution of Potassium Arsenite. *Liquor Potassii Arsenitis.* KAsO_2 .—This solution, popularly known as *Fowler's Solution*, is prepared by boiling together arsenous acid and potassium bicarbonate until solution ensues, flavoring with compound tincture of lavender, and filtering after a week: $\text{As}_2\text{O}_3 + 2\text{KHCO}_3 = 2\text{KAsO}_2 + 2\text{CO}_2 + \text{H}_2\text{O}$.

Chemists are not certain whether the two salts react upon each other, and if they do, whether the orthoarsenite, K_3AsO_3 (or KH_2AsO_3), or the metarsenite, KAsO_2 , is formed. It is probable, though, that the latter is formed. The solution contains 1 per cent. As_2O_3 . It is an agreeable preparation and is used more extensively than any other form of arsenic. Dose is up to 5 minims, diluted. It should never be kept longer than a year, as after that time and often before, it begins to deposit the arsenous oxide,

becoming proportionately weaker, and creating the possibility of poisoning by the careless administration of the deposit. The solution should never be dispensed unless it is perfectly clear and transparent.

Arsenic Iodide. *Arseni Iodidum.* AsI_3 .—The iodide is made by directly combining iodine with arsenic, the union being facilitated by dissolving the iodine in disulphide of carbon and adding the arsenic in powder until the purple color of the iodine disappears, carefully evaporating and crystallizing: $\text{As}_4 + 6\text{I}_2 = 4\text{AsI}_3$. The arsenic iodide thus prepared is in shining, orange-red crystalline scales losing iodine upon exposure to air; volatile, soluble in water and alcohol, gradually decomposing in solution. This compound is rarely used by itself, but is usually combined with iodide of mercury in the preparation of Donovan's Solution.

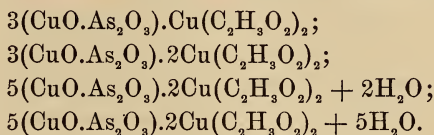
Solution of Arsenic and Mercuric Iodide. *Liquor Arseni et Hydrargyri Iodidi.* $\text{AsI}_3 \cdot \text{HgI}_2$.—One per cent of each of the iodides of arsenic (arsenic) and mercury (mercuric) are dissolved in water and the solution filtered. Probably no chemical reaction takes place. The solution should be of a light straw-color; any deeper color is due to the presence of disengaged iodine, and may be removed by the addition of a globule of mercury, or of a few grains of metallic arsenic, and filtering. This solution, like the other arsenic and mercury preparations, is an alterative. Dose, up to 5 minims, diluted.

Sodium Arsenate. *Sodii Arsenas.* $\text{Na}_2\text{HAsO}_4 \cdot 7\text{H}_2\text{O}$.—See Sodium Salts.

Solution of Sodium Arsenate. *Liquor Sodii Arsenatis.*—A preparation containing one per cent of sodium arsenate dissolved in distilled water.

Paris Green.—This unofficial arsenic compound is an

aceto-arsenite of copper prepared by boiling solution of arsenous acid with solution of acetate of copper. It is of variable composition and may be any of the following:



The last is considered to be the best. It is often termed Vienna or Imperial Green. It should be sold with caution, being exceedingly poisonous.

Antimony Salts.

Occurrence.—Antimony occurs with arsenic chiefly as a sulphide, Sb_2S_3 , and also native. The former when freed from earthy impurities by fusion constitutes the *crude* or *black antimony* of the market.

Preparation.—The metal may be obtained from its ores by heating them with carbonates of potassium or sodium, or with metallic iron, which retains the sulphur, or by roasting and heating with charcoal. The latter is the method usually employed. $2\text{Sb}_2\text{S}_3 + \text{heat} + 9\text{O}_2 = 2\text{Sb}_2\text{O}_3 + 6\text{SO}_2$, $2\text{Sb}_2\text{O}_3 + 2\text{C}_2 = 2\text{Sb}_2 + 2\text{CO}_2 + 2\text{CO}$ ($\text{CO}_2 + \text{CO}$ in variable proportions).

The metal is not used in pharmacy; the purified sulphide furnishes directly or indirectly all of the preparations of the Pharmacopœia. Metallic antimony has numerous applications in the arts; type-metal, pewter, etc., are alloys containing it.

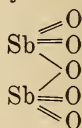
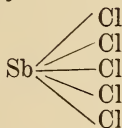
Properties.—Antimony is a bluish-white metal, lustrous, very brittle, and hence easily pulverized; specific gravity, 6.8; volatile at high heat.

It does not oxidize in the air at ordinary temperatures, but when strongly heated it burns, producing the oxide Sb_2O_3 . Dissolved in nitric acid it is oxidized to antimonious acid; hydrochloric acid dissolves it, forming a chloride, SbCl_3 , and free hydrogen. Antimony has chemical properties very similar to those of arsenic; it forms two classes of compounds—the antimonous and antimonic. In the former it is trivalent, SbCl_3 , Sb_2S_3 , Sb_2O_3 , in the latter quinquivalent, Sb_2O_5 , SbCl_5 , Sb_2S_5 .

Graphically, antimonous compounds may be thus expressed:



Antimonic thus,



The atomic weight of antimony is 120; its molecule consists of two atoms, Sb_2 . Nearly all that has been said of the chemical behavior of arsenic is also true of antimony. Antimony forms the base in some compounds, SbCl_3 , and the radical in others, NaSbO_2 sodium antimonite.

The trioxide Sb_2O_3 unites with water to form antimonous acid, $\text{Sb}_2\text{O}_3 + \text{H}_2\text{O} = 2\text{HSbO}_2$, which forms salts called *antimonites* by having its hydrogen replaced by a metal, e.g. KSb_2O_5 , antimonite of potassium.

The pentoxide produces the antimonie acid with water, $\text{Sb}_2\text{O}_5 + \text{H}_2\text{O} = 2\text{HSbO}_3$, from which antimonates are obtained. Both these acids are monobasic. The Sb_2O_5 yields also a dibasic acid with H_2O , thus: $\text{Sb}_2\text{O}_5 + 2\text{H}_2\text{O} = \text{H}_4\text{Sb}_2\text{O}_7$.

Important Reactions (see also Arsenic).—1. H_2S in acidified solutions of antimony salts gives *orange-red* precipitate.

2. ANTIMONY SALTS dissolve in $(\text{NH}_4)_2\text{S}_2$: $2\text{Sb}_2\text{S}_3 + 6(\text{NH}_4)_2\text{S}_2 = 4(\text{NH}_4)_3\text{SbS}_4 + \text{S}_2$.

3. MARSH'S TEST, employed as for arsenic, yields antimonuretted hydrogen, H_3Sb : $\text{Sb}_2\text{O}_3 + 6\text{H}_2 = 2\text{H}_3\text{Sb} + 3\text{H}_2\text{O}$. The H_3Sb ignited and its flame directed upon cold porcelain produces black spots by the deposition of metallic antimony. These spots remain unaffected with solutions of hypochlorites (arsenic spots dissolve and disappear).

If H_3Sb is passed into silver nitrate solution, the antimony is precipitated (Arsenic remains in solution): $\text{H}_3\text{Sb} + 3\text{AgNO}_3 = \text{Ag}_3\text{Sb} + 3\text{HNO}_3$.

4. ANTIMONY is precipitated from its solution as black powder by iron or zinc; copper removes it from solution in form of metallic film.

5. A CONCENTRATED SOLUTION OF THE CHLORIDE poured into water produces a precipitate of the oxychloride: $\text{SbCl}_3 + \text{H}_2\text{O} = \text{SbOCl} + 2\text{HCl}$.

Antimony Sulphide. *Antimonii Sulphidum*. Sb_2S_3 .—The crude ores are fused in such a way that most of the earthy impurities are removed. When cold the sulphide is in steel-gray masses of metallic lustre, or in black lustreless powder, odorless, insoluble, tasteless. The U. S. P. directs it to be "as free from arsenic as possible." It is used mainly to prepare the purified sulphide in pharmacy, and as an alterative in veterinary practice.

Purified Antimony Sulphide. *Antimonii Sulphidum Purificatum.* Sb_2S_3 .—The above-described sulphide, finely powdered, is macerated for several days with ammonia-water, thoroughly washed and dried. The ammonia-water dissolves the sulphide of arsenic, which is associated with the antimony in the ores, leaving the antimony sulphide as residue (see Second Reaction of Arsenic Salts).

The sulphide thus purified is a dark-gray powder, odorless, tasteless, insoluble in water, soluble in hot hydrochloric acid with the evolution of H_2S gas. It should not contain more than mere traces of arsenic. The orange-red precipitate obtained by passing H_2S gas into solutions of antimony has the same composition, Sb_2S_3 . It is changed to the black variety by heating.

It is not used internally, but enters into the preparation of the Sulphurated Antimony.

Sulphurated Antimony. *Antimonium Sulphuratum.* Sb_2S_3 .- Sb_2O_3 .—This preparation, sometimes called *Kermes Mineral*, is obtained by boiling the purified antimony sulphide with soda solution and adding sulphuric acid: $6\text{NaOH} + \text{Sb}_2\text{S}_3 = \text{Na}_3\text{SbO}_3$ (sodium antimonite) + Na_3SbS_3 (sodium sulph-antimonite) + $3\text{H}_2\text{O}$. To the hot solution of the latter two compounds H_2SO_4 is added which precipitates the sulphide and oxide: $2\text{Na}_3\text{SbO}_3 + 3\text{Na}_3\text{SbS}_3 + 6\text{H}_2\text{SO}_4 = 6\text{Na}_2\text{SO}_4 + \text{Sb}_2\text{S}_3 + \text{Sb}_2\text{O}_3 + 3\text{H}_2\text{S} + 3\text{H}_2\text{O}$. The precipitate, consisting almost wholly of the Sb_2S_3 , is thoroughly washed to remove the sodium sulphate which remained in solution and dried. It is a reddish-brown amorphous powder, odorless, tasteless, insoluble. It should be free from sulphate, ascertained with BaCl_2 . (There are many varieties of sulphurated antimony; the above is the official one.)

It is emetic and alterative in doses up to 0.330 Gm. Usually given in pill form.

Compound Pills of Antimony. *Pilulæ Antimonii Compositæ*, Plummer's Pills.—Pills containing each 0.040 Gm. of sulphurated antimony, 0.040 Gm. of calomel and 0.080 Gm. of guaiac made into mass with mucilage of tragacanth. No chemical reaction takes place between the ingredients. Given in skin diseases.

Antimony Oxide. *Antimonii Oxidum*. Sb_2O_3 .—Obtained by dissolving Sb_2S_3 in hydrochloric acid, pouring the solution into much water and treating the resulting precipitate with ammonia-water: $\text{Sb}_2\text{S}_3 + 6\text{HCl} = 2\text{SbCl}_3$ (antimonous chloride) $+ 3\text{H}_2\text{S}$. $10\text{SbCl}_3 + 15\text{H}_2\text{O} = 5\text{Sb}_2\text{O}_3$ (antimonous oxide) $+ 30\text{HCl}$. Not all of the SbCl_3 is converted into the oxide, but some becomes the oxychloride, SbClO : $2\text{SbCl}_3 + 2\text{H}_2\text{O} = 2\text{SbClO} + 4\text{HCl}$. The SbClO is converted into the oxide by having its chlorine removed with NH_4OH : $2\text{SbClO} + 2\text{NH}_3 + \text{H}_2\text{O} = \text{Sb}_2\text{O}_3 + 2\text{NH}_4\text{Cl}$. Oxide of antimony is a whitish powder, permanent in the air, odorless, tasteless, almost insoluble in water, insoluble in alcohol, soluble in hydrochloric or nitric or tartaric acids. It should be free from arsenic, chloride and sulphate. It enters into the preparation of antimonial powder and tartar emetic.

Antimonial Powder. *Pulvis Antimonialis*. $\text{Sb}_2\text{O}_3 + \text{Ca}_3(\text{PO}_4)_2$.—A powder consisting of 33 per cent oxide of antimony and 67 per cent precipitated phosphate of calcium, and given as a diaphoretic in doses of 0.20 to 0.330 Gm.

Antimony and Potassium Tartrate. *Antimonii et Potassii Tartras*. $2\text{KSbOC}_4\text{H}_4\text{O}_6 \cdot \text{H}_2\text{O}$, Tartar Emetic.—Made by boiling together antimony oxide and potassium bitartrate and allowing the solution to crystallize: $2\text{KHC}_4\text{H}_4\text{O}_6 +$

$\text{Sb}_2\text{O}_3 = 2\text{KSbOC}_4\text{H}_4\text{O}_6 + \text{H}_2\text{O}$. The hydrogen in the base of the acid tartrate of potassium is replaced by SbO. Tartar emetic is in form of granular powder or as small transparent crystals, of a sweetish taste, soluble in water, insoluble in alcohol. It should be free from arsenic.

This is the most important of the antimony compounds, being used extensively, as its name indicates, as an emetic. All the antimonial preparations have more or less emetic properties, a fact which accounts for the absence of cases of poisoning by antimony. The antidote would be tannin, which produces the insoluble tannate of antimony.

Wine of Antimony. *Vinum Antimonii*. $\text{KSbOC}_4\text{H}_4\text{O}_6$.—A preparation made by dissolving 4 parts of tartar emetic in 1000 parts white wine containing 150 parts of alcohol. It is usually added to expectorant mixtures. Dose, 5 to 10 minims.

Bismuth Salts.

Occurrence.—Bismuth occurs in the metallic state, and often as sulphide associated with traces or more of arsenic.

Properties.—Bismuth is a grayish-white metal with a pinkish tint; it is crystalline, brittle, brilliant; specific gravity 9.8, and has some resemblance to metallic antimony. Chemically it is closely allied to antimony. It is usually trivalent, sometimes quinquivalent; atomic weight 208.9. There are only four of its compounds official, and beyond the employment in medicine it is little used.

Analytical Reactions.—1. WATER precipitates from acidulated bismuth solution white *basic* salts which contain more or less of the acid radical in proportion as little or much water is used: $\text{BiCl}_3 + \text{H}_2\text{O} = \text{BiOCl} + 2\text{HCl}$ (see Antimony).

2. ALKALINE HYDROXIDES precipitate bismuth hydroxide: $\text{BiCl}_3 + 3\text{KOH} = \text{Bi}(\text{OH})_3 + 3\text{KCl}$.

3. HYDROSULPHURIC ACID AND SULPHIDES produce black precipitate of Bi_2S_3 which is insoluble in alkali sulphides, affording a means of separating bismuth from arsenic and antimony: $2\text{BiCl}_3 + 3\text{H}_2\text{S} = \text{Bi}_2\text{S}_3 + 6\text{HCl}$.

Bismuth Subnitrate. *Bismuthi Subnitratis.* $\text{BiONO}_3 \cdot \text{H}_2\text{O}$.—Prepared in a necessarily circumstantial way by dissolving bismuth in diluted nitric acid, diluting the resulting solution with a prescribed quantity of water, and allowing to stand for a day, then filtering. The filtered solution is then poured into a solution of sodium carbonate during constant stirring, the precipitate drained and washed and dissolved in nitric acid. The nitric-acid solution is again allowed to stand for a day, then again diluted with a prescribed quantity of water, and some water of ammonia added, and the precipitate thoroughly washed and dried.

Metallic bismuth as a rule contains arsenic, and in the above process provision is made for its separation. The *rationale* of the process is as follows: The bismuth dissolves to bismuth nitrate in nitric acid, and some of the arsenic becomes arsenic acid: $\text{Bi}_2 + 8\text{HNO}_3 = 2\text{Bi}(\text{NO}_3)_3$, bismuth nitrate (normal), $+ \text{N}_2\text{O}_2 + 4\text{H}_2\text{O}$, $3\text{As}_4 + 20\text{HNO}_3 + 8\text{H}_2\text{O} = 12\text{H}_3\text{AsO}_4$, arsenic acid, $+ 10\text{N}_2\text{O}_2$.

The bismuth nitrate and the arsenic acid react upon each other, producing bismuth arsenate, $\text{Bi}(\text{NO}_3)_3 + \text{H}_3\text{AsO}_4 = \text{BiAsO}_4 + 3\text{HNO}_3$. Both of these salts when their solution is diluted with water begin to separate and deposit, but the arsenate needs less water to cause it to separate than the nitrate does, and hence just enough water is added to the solution to cause the arsenate, but not the nitrate, to deposit. When the arsenate has deposited, after a day's

standing, it is removed by filtration. Any arsenate which may not have deposited when the sodium carbonate is added becomes the very soluble arsenate of sodium, which remains in solution while the bismuth is precipitated as the insoluble subcarbonate: $\text{Na}_2\text{CO}_3 + \text{H}_3\text{AsO}_4 = \text{Na}_2\text{HAsO}_4$, arsenate of sodium, $+ \text{H}_2\text{O} + \text{CO}_2$. $3\text{Na}_2\text{CO}_3 + 2\text{Bi}(\text{NO}_3)_3 = \text{Bi}_2\text{O}_2\text{CO}_3$, subcarbonate bismuth, $+ 6\text{NaNO}_3 + 2\text{CO}_2$. The subcarbonate is washed and the arsenate of sodium removed in the washings. As a further precaution to insure the entire absence of arsenic, the subcarbonate is again dissolved in nitric acid, and the solution again diluted to a point at which any arsenic still present will precipitate according to above equation.

After some time the normal nitrate is converted into the subnitrate or oxynitrate by further dilution with water and the neutralization of the nitric acid present with ammonia water: $\text{Bi}(\text{NO}_3)_3 + \text{H}_2\text{O} = \text{BiONO}_3 + 2\text{HNO}_3$, $\text{HNO}_3 + \text{NH}_4\text{OH} = \text{NH}_4\text{NO}_3 + \text{H}_2\text{O}$. The subnitrate of bismuth thus prepared is a heavy white powder permanent in the air, odorless, almost tasteless and insoluble. It should be free from carbonate: effervescence with acids would indicate its presence.

Subnitrate of bismuth is largely used in stomach troubles as sedative and tonic. Dose up to 0.66 Gm.

Bismuth Subcarbonate. *Bismuthi Subcarbonas.* $\text{Bi}_2\text{O}_2\text{CO}_3 \cdot \text{H}_2\text{O}$.—The subcarbonate is made as above shown in the making of the subnitrate. (See the preceding for equations and rationale.)

It is a pale yellowish-white powder, permanent in the air, odorless, tasteless and insoluble. It may contain more impurities than the subnitrate because in the latter the process of purification is further carried on, especially in

the elimination of arsenic. The impurities which may be present are insoluble foreign salts, alkalies and alkaline earths, lead, copper, chloride and sulphate, and traces of arsenic and antimony.

Its use is similar to that of the subnitrate, and the dose the same.

Bismuth Citrate. *Bismuthi Citras.* $\text{BiC}_6\text{H}_5\text{O}_7$.—This salt is prepared by boiling bismuth subnitrate with citric acid and water for 15 or 20 minutes, or until a drop of the mixture dissolves in ammonia-water. Then the suspended matter is allowed to subside, the solution decanted and the deposit washed with water on a filter until the washings are tasteless. Thus obtained it is usually an amorphous powder without taste or odor and insoluble in water, but soluble in ammonia-water and in solutions of the citrates of the alkalies. It is official mainly for its employment in the bismuth and ammonium citrate. Incomplete washing in its preparation usually results in the presence of nitrate in the salt. The usual test for nitrates should be applied.

Bismuth and Ammonium Citrate. *Bismuthi et Ammonii Citras.*—The introduction of this salt into medicine was mainly to obtain a soluble salt of bismuth. It is not a definite compound, but it is soluble. Its preparation consists in dissolving bismuth citrate in ammonia-water, evaporating (which also dispels all excess of ammonia) to a syrupy consistency, and spreading on plates of glass or porcelain, so that when dry the salt may be obtained in scales. The product should be kept in small well-stoppered bottles protected from light. The small shining, pearly scales are odorless and have a slightly metallic taste. They become

opaque and less soluble if exposed to the air, hence the necessity of keeping them in well-closed vessels.

The presence of ammonia in the compound can easily be detected by boiling with soda or potassa solution, which liberates the ammonia, which may be detected by its odor.

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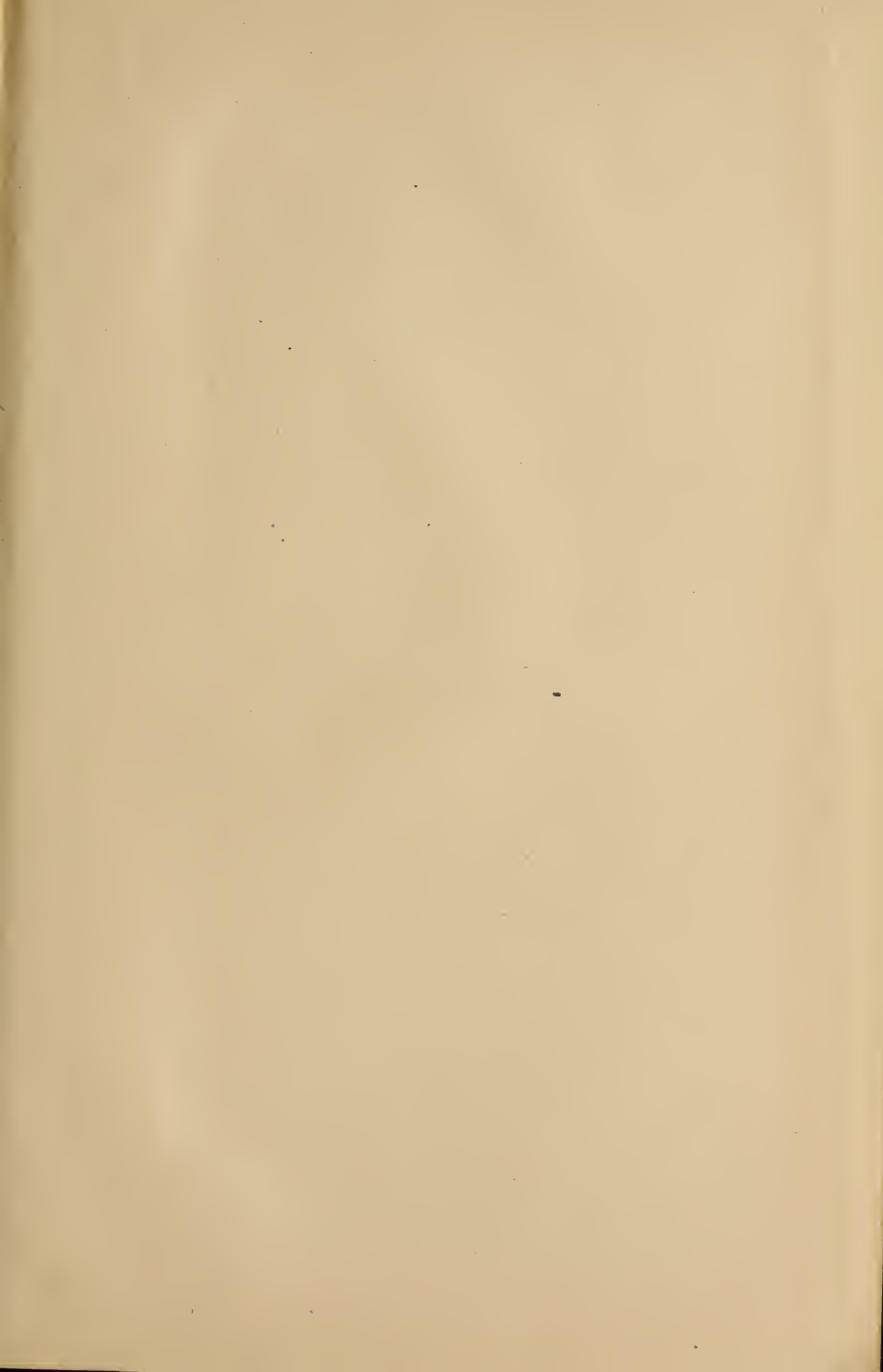
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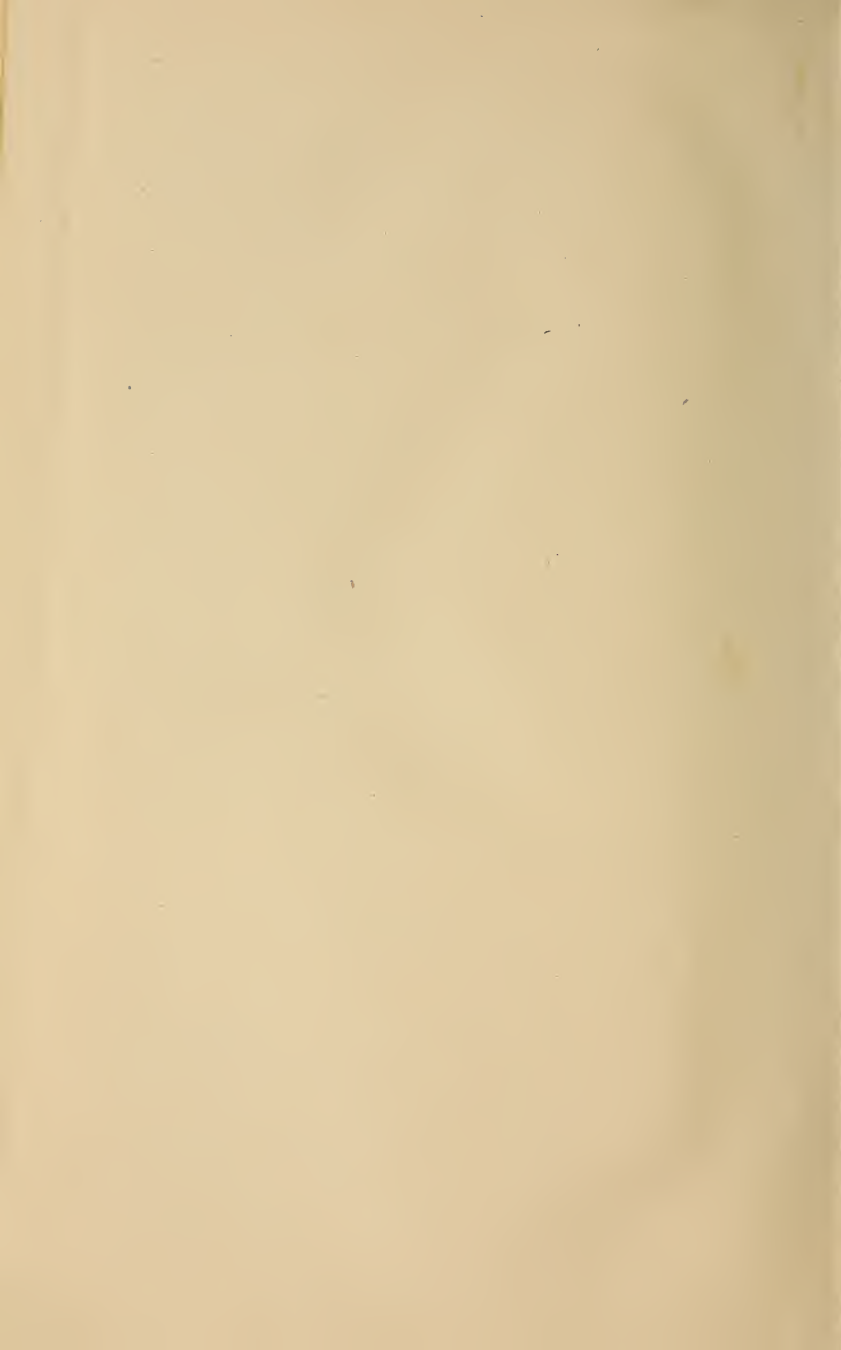
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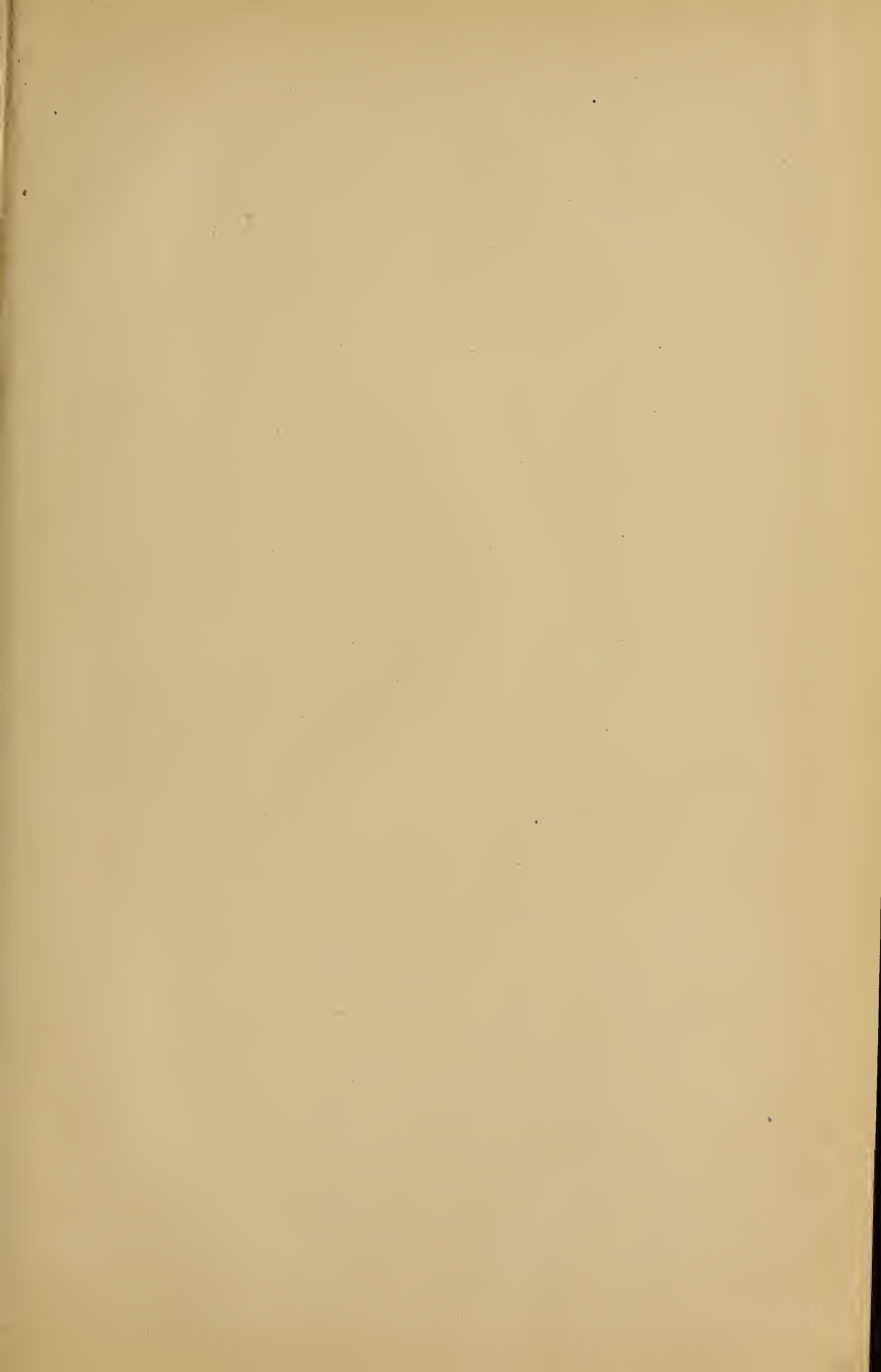
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